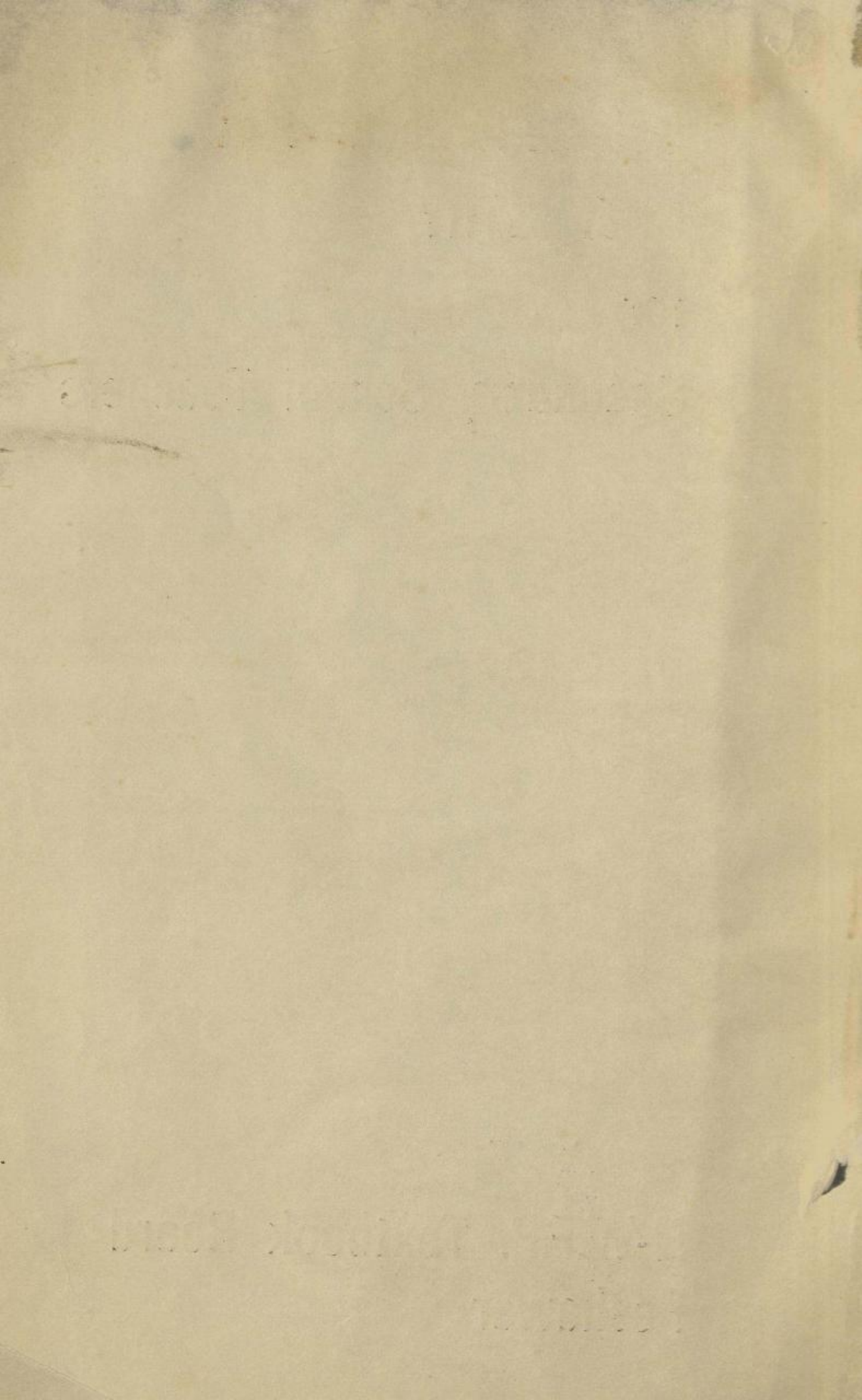


1986

Selected Modules in
CHEMISTRY
for
Secondary School Teachers



N-W.F.P. Textbook Board
Peshawar

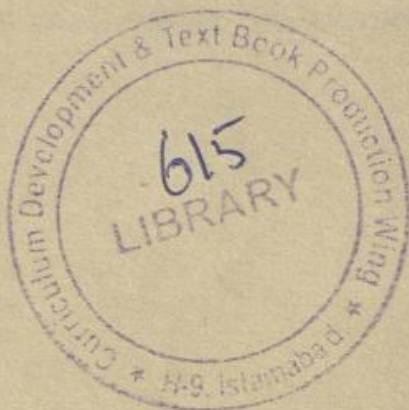


ACKNOWLEDGEMENT

I am thankful to the participants, the local resource persons and British team for helping us in producing this monograph on some topics from Secondary School Chemistry generally considered as difficult by our teachers and students community. This Monograph is the by-product of a teacher training workshop held during Summer 1983 and was produced besides many other teacher training activities in a record time of two weeks. My thanks are therefore due to the entire team who took extra pains and produced various units by working even after routine working hours.

Despite the fact that this is the first attempt of its kind and that there is much room for improvement in it, it is hoped that it will help the students and the teachers in understanding the topics included in it. Work pertaining to the compilation and editing of the report was jointly done by Dr. M. Jaffar, Associate Professor, Department of Chemistry, Quaid-i-Azam University, Islamabad and Dr. R. A. Siyal, Assistant Educational Adviser, Ministry of Education, Islamabad.

In the forthcoming seminars we intend to produce similar monographs on selected topics so that it could be used by our Secondary School Chemistry Teachers, parents and students to gain meaningful in-sight in Chemistry.



Abdullah Khadim Hussain
Joint Educational Adviser
(Curriculum Wing),
Ministry of Education,
Islamabad.

INTRODUCTION

The teaching of Science in our country leaves much to be desired for a variety of reasons. First properly qualified teachers are not available. Secondly laboratory facilities are in-adequate. Thirdly, there are no back-up facilities to make up for these shortcomings. In order to up grade the competencies of teachers, Ministry of Education regularly organizes courses for Master Teacher Trainers who in turn provide services for the training of teachers at the provincial level. This is a continuous process, particularly because the knowledge of the teachers needs up-dating both in terms of subject matter as well as methodology. This year the Federal Ministry of Education organized Summer Seminars in collaboration with the British Council for Training of Master Teacher Trainers in the subjects of Physics, Chemistry, Biology, Mathematics and English at Secondary level. In each subject about 20 senior teachers from various Teacher Training Organizations were trained.

In order to ensure that the modules generated during the training with the help of the teacher trainers reach the teachers, parents and students, it was decided to identify difficult or important topics and develop materials which would facilitate learning of these topics. This monograph is the outcome of the collective effort of all those involved in the Summer Seminars in Chemistry. Our thanks are particularly due to the Pakistani and British resource persons who guided the deliberations of the Seminar and made the writing of this monograph possible. We hope that it would adequately serve the purpose for which it has been developed.

Saeed Ahmad Qureshi
Secretary,
Ministry of Education
Government of Pakistan
Islamabad.

A UNIT

ON

ELECTROCHEMISTRY

PREPARED BY

MIAN BASHIR AHMAD MUSHTAQ

AND

DR. MOHAMMAD AFZAL BHATTI

IN A

SUMMAR SEMINAR IN CHEMISTRY

ORGANIZED BY

GOVERNMENT OF PAKISTAN

MINISTRY OF EDUCATION

(CURRICULUM WING)

ISLAMABAD

A UNIT
ON
ELECTROCHEMISTRY

A. BRIEF INTRODUCTION TO THE TOPIC

You must have observed that whenever a chemical reaction takes place, energy is either evolved or absorbed. This energy is usually in the form of heat, light or sound and can be converted to mechanical or electrical energy. A chemical reaction in which energy is released in the form of electricity is called an electrochemical reaction, e.g., getting light by means of an electric torch. The field of chemistry which deals with the study of chemical reactions involving evolution or absorption of electrical energy is called electrochemistry.

B. LIST OF CONTENTS/CONCEPTS

This unit includes the following sub-topics/concepts :

- a) **Electrical Con-**
 - i) Conductors/Non conductors or insulators;
 - ii) Electrolytes/Non electrolytes;
 - iii) Conductance in metals and electrolytes.
- b) **Electrolysis** i) Process and products of electrolysis and Faraday's laws of electrolysis; (ii) weight relationship in electrochemical reactions.
- c) Production of electric current by chemical reactions, Voltaic Cell.

C. OBJECTIVES

- At the end of the unit the pupil should be able to :
1. Define an electrochemical process, electrochemistry, electrolytes, electric conductance, electrolysis, Voltaic or Galvanic Cell, electrode potential and electroplating.
 2. Classify substances as conductors/non conductors/ electrolytes/non electrolytes.
 3. Identify conductors/non conductors and electrolytes/non electrolytes through experiments.
 4. Have knowledge of the process and products of electrolysis.
 5. Demonstrate Faraday's laws of electrolysis by simple experiments.
 6. Solve mathematical problems concerned with gram equivalent of a metal or weight of a metal deposited on an electrode from an electrolyte.
 7. Construct and explain the working of a Voltaic or Galvanic Cell.
 8. Demonstrate that chemical energy can be converted to electrical energy.

D. INSTRUCTIONAL ACTIVITIES

ELECTRIC CONDUCTANCE

a) Identifying conductors and non conductors

Experiment 1...Conduction of electricity through elements.

Take pieces of elements like carbon (a rod of graphite), copper, lead, zinc and sulphurand fit each piece turn by turn (as shown in Figure 1) to the two clips and see if the bulb glows or not.

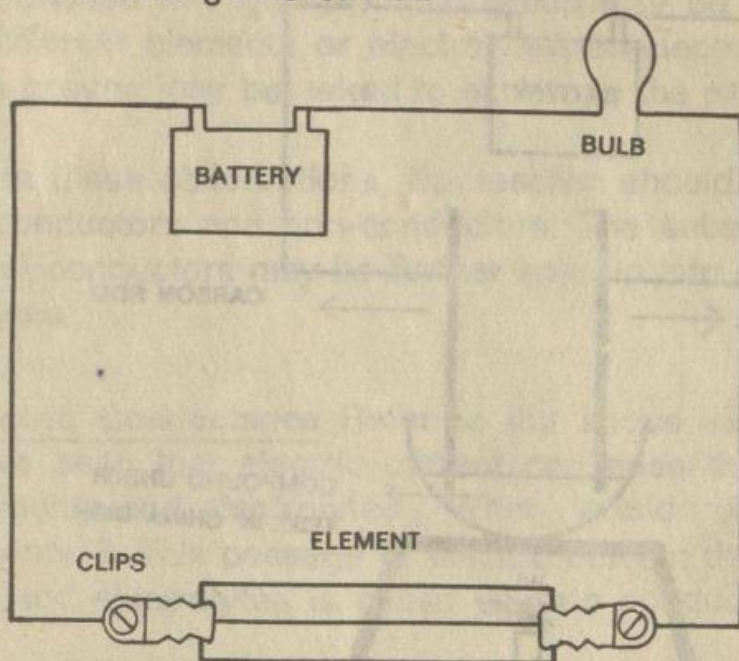


Fig. 1

The teacher should encourage the pupils by asking the following questions about the experiment.

1. Make a list of elements that conduct or do not conduct electricity.
2. Do all metals in your list conduct electricity?
3. Do all non-metals conduct electricity?
4. What inference do you draw from these observations?

b) Electrolytes/Non electrolytes

Experiment. 2 Conduction of electricity by compounds.

1. Set up the apparatus as shown in Figure 2.

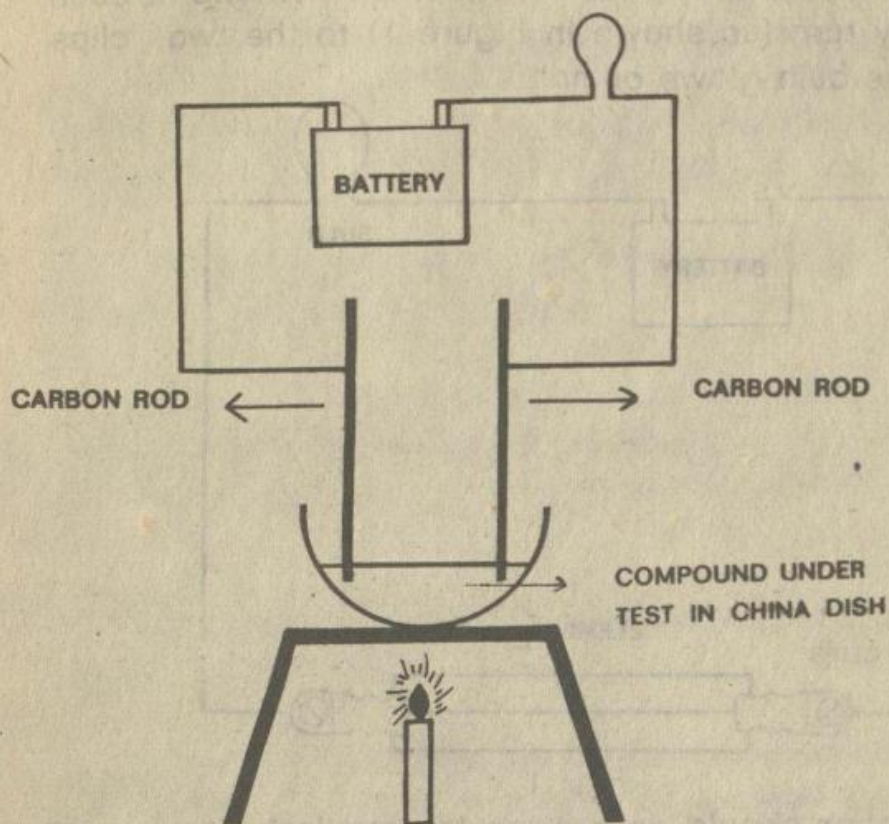


Fig. 2

2. Use fused compounds or compounds dissolved in water and see if the bulb glows (compounds like fused Pb Br_2 and aqueous CuSO_4 may be used).
3. What do you infer from it? (Ask the students).
4. What do you see at the anode and cathode? (Derive definition of electrolytes from student's responses).
5. Now use fused wax, distilled water and kerosene in turn in the china dish, to demonstrate the electrical behaviour of non electrolytes.

Students' experiments Let the students perform the above experiments by using different conducting and non-conducting elements and electrolytes. For this purpose the class may be divided into groups, each group may be asked to use a different elements or electrolyte/non-electrolyte and then the groups may be asked to exchange the experiments.

From these observations, the teacher should develop a list of conductors and non-conductors. The substances in the list of conductors may be further split up into metals and electrolytes.

c) **Electric Conductance** Refer to the above experiments. You have seen that electric current can pass through certain elements and electrolytes. What would you call this phenomenon? This passage of electric current through conductors and electrolytes is called electric conductance.

d) **Electrolysis**

Teacher's demonstration

Experiment 3 (1) Set up the apparatus as shown in Figure 3.

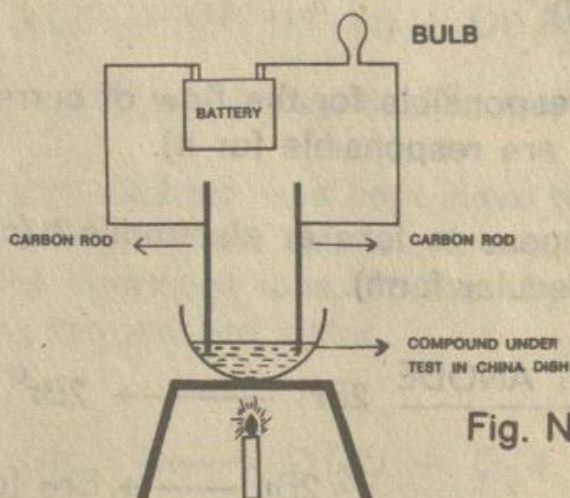


Fig. No. 3

- Place some fused Pb Br_2 crystals in a crucible and switch on the current.
- Note the products at the cathode (negative electrode) and anode (positive electrode).
- Tabulate the results as below :

Substance Electrolysed	Observation at cathode	Product at cathode	Observation at anode	Product at anode
Lead Bromide	Greybead of metal forms	Lead	Dark red brown gas forms	Bromine

- Lead the pupils in discussion with questions such as those which follow :

- What happens to Pb Br_2 ? (It is decomposed into ions; $\text{Pb Br}_2 \longrightarrow \text{Pb}^{2+} + 2\text{Br}^-$).

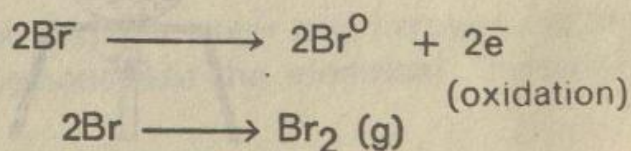
- Why do lead ions move to cathode ? (They are positively charged).

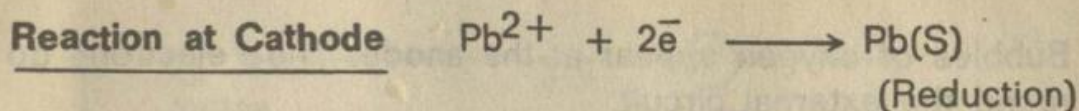
- Why do bromide ions move to anode ? (They are negatively charged).

- What is responsible for the flow of current through the solution (ions are responsible for it).

- What happens to ions at electrodes ? (converted into atomic or molecular form).

REACTION AT ANODE

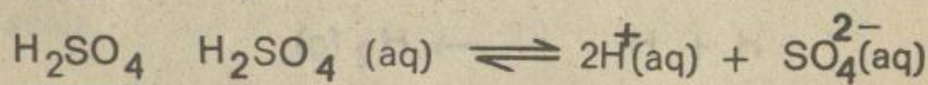
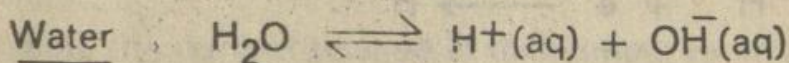




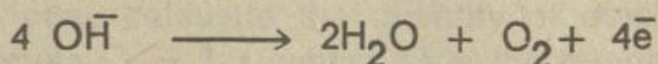
6. What are the ions collected at the anode called ? (anions).
7. What are ions collected at the cathode called ? (cations).
8. Repeat the experiment by using aqueous solution of PbBr_2
9. Note that the above process is called electrolysis.
10. How would you define it ? (When electricity passes through fused or aqueous solution of electrolytes, chemical changes occur at the electrodes where new substances are produced. This process of causing chemical reactions by passing electric current through electrolytes is called electrolysis).

Students's experiments

1. Repeat the above experiment by taking acidified water in the beaker and note the products at the electrodes as below :

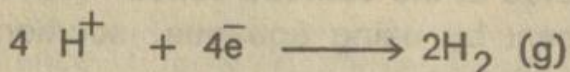


Hydroxide ions and sulphate ions both move to anode. But it requires more energy for the sulphate ions to give up their electrons than the hydroxide ions. So the hydroxide ions break up forming oxygen and water.



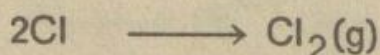
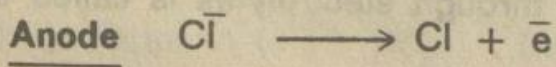
Bubbles of oxygen appear at the anode. The electrons go round the external circuit.

There are only hydrogen ions to be attracted towards the cathode and they accept electrons to form hydrogen molecules.

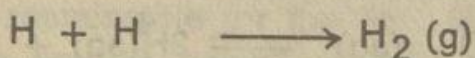
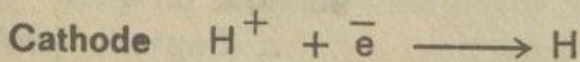


So bubbles of hydrogen appear at the cathode.

2. Repeat the same experiment by taking aqueous ZnCl_2 in the beaker and note the products at the electrodes as given below :



So chlorine with greenish yellow colour and specific odour appears at the anode.



e) Faraday's laws of Electrolysis

Experiment 4. (1) Set up the apparatus as shown in Figure 4.

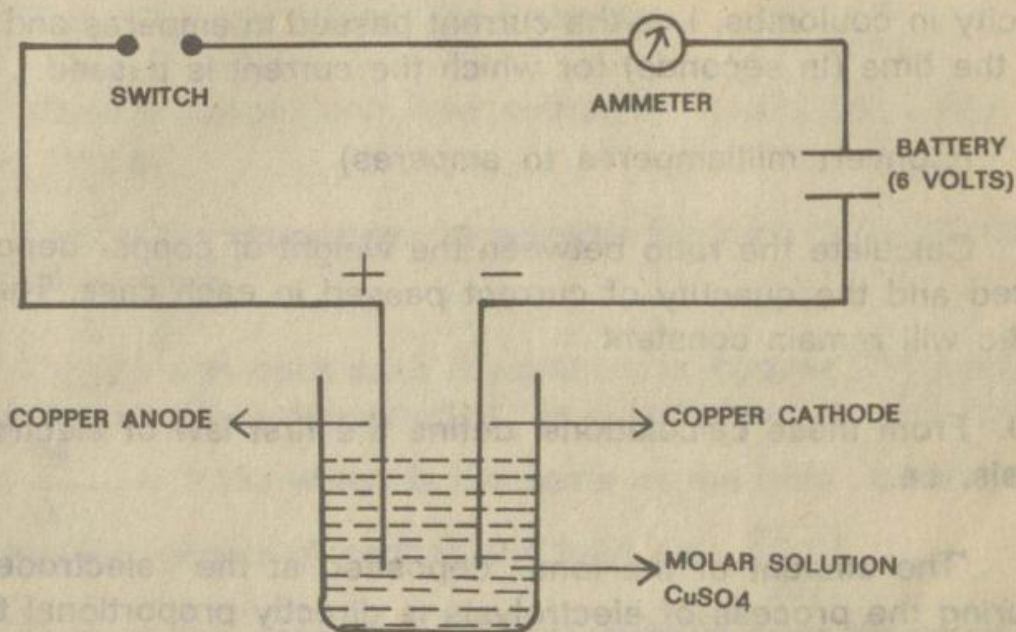


Fig. 4

2. Pour solution of M CuSO_4 into the beaker.
3. Clean copper electrodes. Weigh one electrode and connect it to the negative pole and connect the other one to the positive pole.
4. Allow the current to pass for 5 minutes. Use a stop-watch for noting the time.
5. Note the current passed in milli-amperes onto the ammeter scale.
6. Remove the cathode. Dry it and weigh it to find out the amount of copper deposited on it.
7. Repeat the experiment by passing the current for 10, 15 and 20 minutes and find out the weight of copper deposited.
8. In each case calculate the quantity of electricity used from the formula $Q = I \cdot t$, where Q is the quantity of elec-

tricity in coulombs, I , is the current passed in amperes and t is the time (in seconds) for which the current is passed.

(Convert milliamperes to amperes).

9. Calculate the ratio between the weight of copper deposited and the quantity of current passed in each case. This ratio will remain constant.

10. From these calculations, define the first law of electrolysis, i.e.,

'The amount of the ions deposited at the electrodes during the process of electrolysis is directly proportional to the quantity of electricity passed.'

Experiment 5 (1) Set up the apparatus as shown in Figure 5. using thoroughly cleaned electrodes of copper and lead. Weigh the electrode connected to negative pole of the battery.

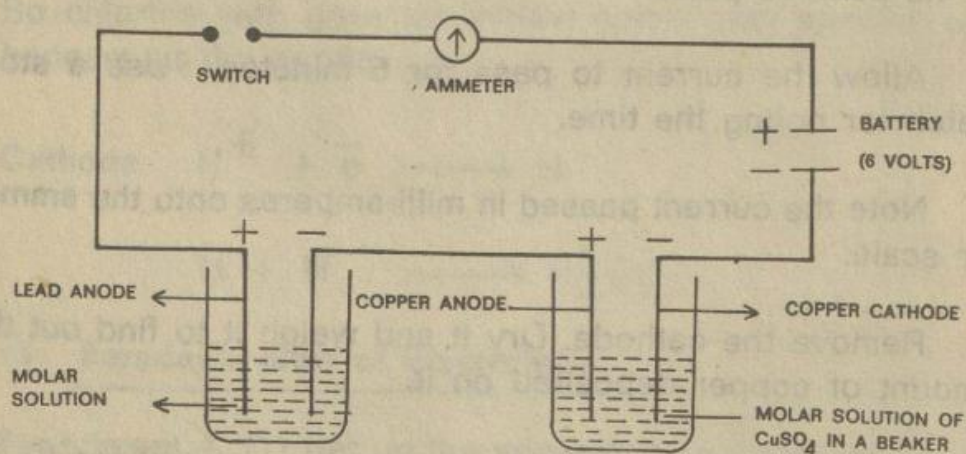


Fig. 5

2. Pour solutions of M CuSO_4 and $\text{M Pb(NO}_3)_2$ in the respective beakers.

3. Switch on the current for 5 minutes.
4. Remove copper and lead cathodes and weigh them after drying.
5. Repeat the experiment by passing the current for 10, 15 and 20 minutes.
6. Calculate in each case the amount of copper (W_1) and that of lead (W_2) deposited at the electrodes. You will find that $\frac{W_1}{W_2} = 0.153$ which is the same as the ratio between equivalent weights of copper and lead, i.e. $\frac{31.75}{207}$.
7. Deduce the second law of electrolysis from these observations, i.e., if the same quantity of electricity is passed through different electrolytes, the mass of the substances deposited at the electrodes by a given amount of electricity is directly proportional to their chemical equivalents.

(f) Application of laws of Electrolysis

These laws can be used to solve related numerical problems through following relationships :

First law $W \propto Q$ where W is the amount of ions deposited and Q is the quantity of electricity passed.

$$\text{So, } W = \text{constant} \times Q$$

$$\text{or } \frac{W}{Q} = \text{constant}$$

Second law $\frac{W_1}{W_2} = \frac{\text{gram equivalent of first element}}{\text{gram equivalent of second element}}$

where W_1 and W_2 represent the quantities of ions of the first and second elements deposited respectively from two electrolytes when the same quantity of electricity is passed through them.

ELECTRODE POTENTIAL

In the textbook there is no explanation available of the term "potential difference" when a chemical cell is described. So the teacher should explain first the term 'electrode potential'.

Experiment. 6 Refer to Figure 6.

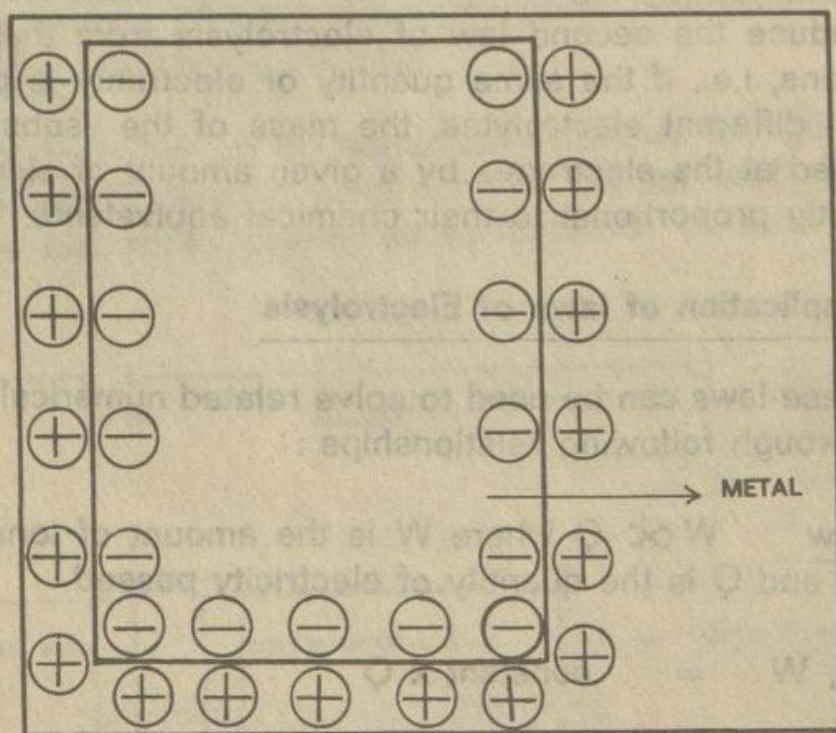


Fig. 6

When a metal is placed in water or in an ionic solution, the metal tends to dissolve. The surface layer of atoms changes to form ions which enter the solution, leaving electrons behind on the surface of the metal. Two layers of

opposite charges are formed. These layers cause a potential difference between the negative charges on the metal and the positively charged ions in the solution. The greater the number of the charged ions in the solution, the greater the tendency for the metal to ionize, the greater will be the potential difference. The potential difference between the metal and the solution of the metal ions is called the metal's electrode-potential. The electrode potential for different metals will depend on the tendency of the metal to form ions. This is explained by the following experiment.

Experiment. 7 1. Set up the apparatus as shown in Figure 7 using a strip of one of the metals listed in (3) below.

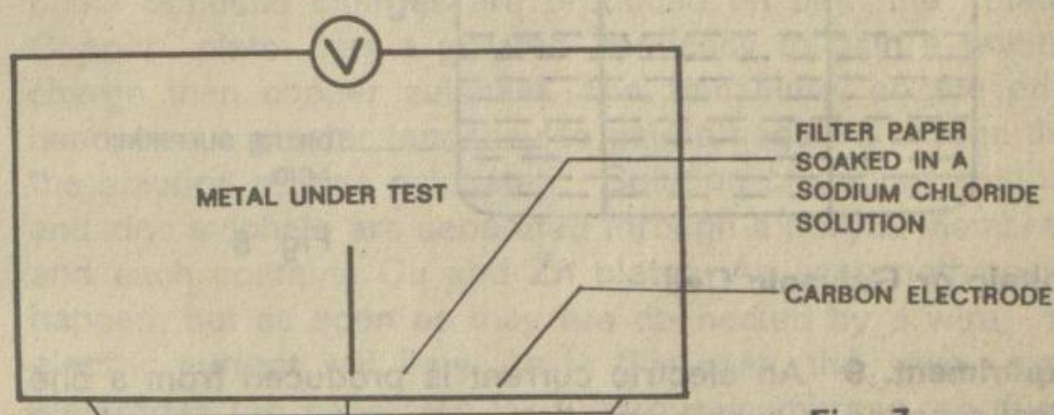


Fig. 7

2. Complete the circuit and record the readings on the voltmeter.
3. Repeat using other metals, e.g., magnesium, aluminum, zinc, tin, lead and silver.
4. Tabulate your results.
5. Compare the behaviour of each metal with the rest in series.

Experiment 8

If two metals, e.g. Cu and Zn, are placed in a beaker containing dilute H_2SO_4 and they are connected with a voltmeter, there will be deflection observed at the voltmeter showing flow of electric current and this flow is due to different electrode potentials of the metals.

The difference of electrode potential of the two metals placed in the same solution is the potential difference of the two metals.

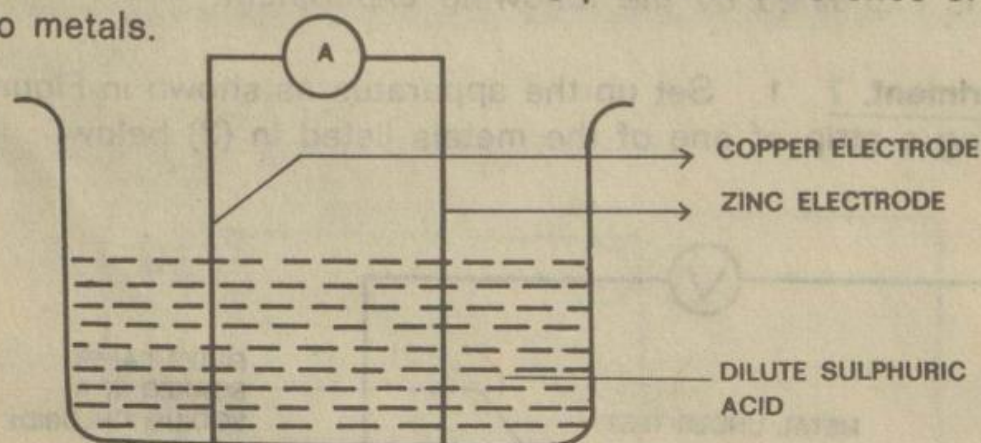


Fig. 8

Voltaic or Galvanic Cell

Experiment. 9 An electric current is produced from a chemical reaction in such a way that a circuit is set up as shown in Figure 8. In this experiment a beaker two third full of a solution of 0.5 M MgSO_4 contains Mg strip and copper foil dipped in it.

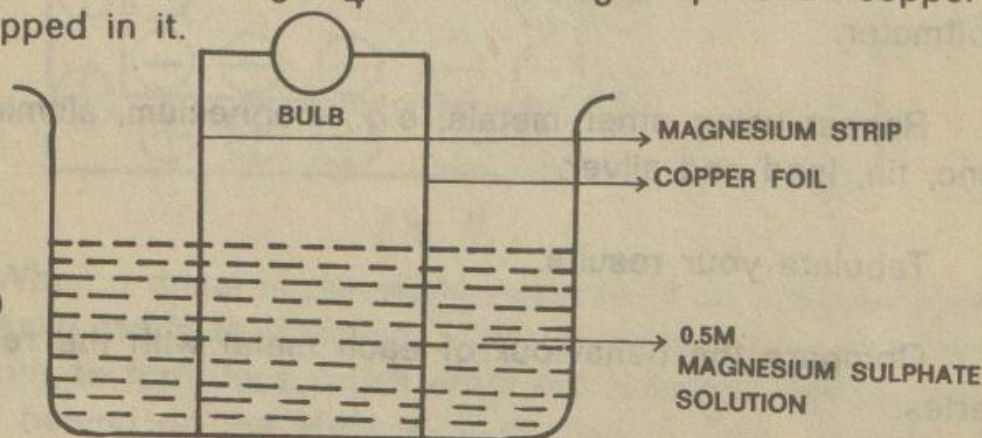
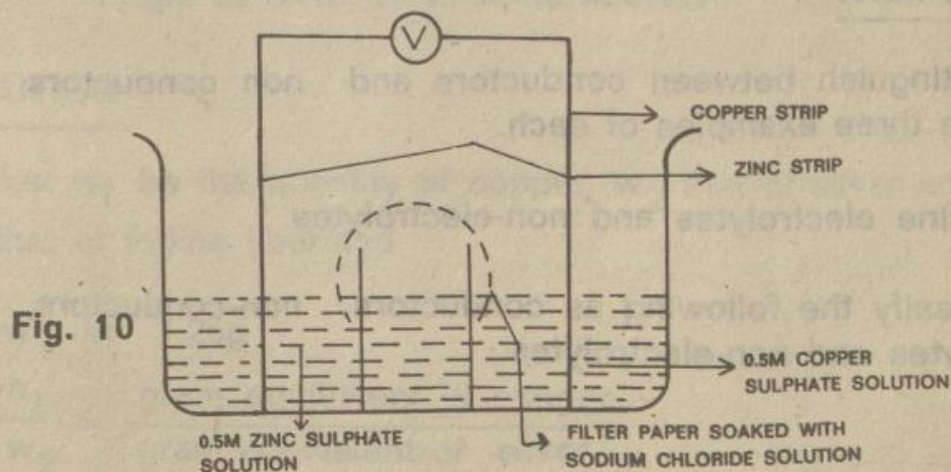


Fig. 9

Connect these metals with the bulb using the connecting wires and crocodile clips. A reaction will take place quite rapidly and the bulb will glow, indicating the flow of electric current in the circuit. The arrangement of two different metals in an electrolyte as in the above experiment is known as a galvanic couple, and the whole system supplying an electric current is called a chemical cell.

Such a cell can work effectively only if the metals are widely separated in the electro-chemical series, e.g., copper and magnesium.

When a copper plate is placed in a solution of copper sulphate and zinc plate is placed in a solution of zinc sulphate opposite charges are produced on both the plates. Copper plate has a greater tendency to gain a positive charge than copper sulphate. The zinc plate, on the other hand, has a greater tendency to gain a negative charge than the solution of zinc sulphate. Solution of copper sulphate and zinc sulphate are separated through a porous membrane and each contains Cu and Zn plates. As such nothing will happen, but as soon as they are connected by a wire, the electric current will flow. As in this case, the two metal electrodes are separated by a porous membrane, so this is the best example of voltaic or Galvanic Cell.



If the teacher has no porous membrane in his laboratory he can use the salt bridge. Two separate beakers contain bridge to complete the circuit as shown in Figure 9. Salt bridge may be obtained by soaking a folded filter paper in a concentrated solution of NaCl.

USING ELECTROCHEMISTRY FOR ELECTROPLATING

This is an opportunity to show that Electro-chemistry has practical applications. Chromium plating on cars and house-hold articles is familiar to most boys and girls, and a teacher may start by explaining to the class that the same principles that they have studied in the previous sections are used in industry for electroplating. They may do a simple experiment on electroplating without a battery by dipping their knives or other steel articles into a copper sulphate solution. The blades become coated with copper. Explain that most of the metals do not get electroplated in this way, as electrical energy is needed to make the reaction go. Let the pupils do their own electroplating experiment as given in the textbook.

ASSESSMENT

1. Distinguish between conductors and non conductors and give three examples of each.
2. Define electrolytes and non-electrolytes.
3. Classify the following as conductors, non-conductors, electrolytes and non-electrolytes :

Wax, ethanol, sodium chloride, sugar solution, steel, plastic, brass, dilute hydrochloric acid, wood, sodium, naphthalene solution and molten lead chloride.

4. Draw a labelled diagram of an electrolysis circuit containing a bulb, a switch, a battery, a carbon cathode, a copper anode and a beaker containing an electrolyte.

5. What does the term electrolysis mean? Draw the apparatus you would use to electrolyse molten magnesium chloride. What would be the products? Write the two electrode reactions and say which of them is an oxidation and which of them is reduction.

6. Define electric conductance. On what factors does it depend?

7. Define Faraday's laws of electrolysis. Apply the laws to solve the following problems:

- 1) Three cells connected in series contain solutions of copper sulphate, silver nitrate and potassium iodide respectively. On passing a fixed amount of electricity, 1.25 g of copper were deposited. Calculate weight of silver and iodine liberated.

SOLUTION

Let w_1 be the quantity of copper, w_2 that of silver and w_3 that of iodine liberated.

$$w_1 = 1.25\text{g}$$

$$\frac{w_1}{w_2} = \frac{\text{gram equivalent of copper}}{\text{gram equivalent of silver}}$$

$$\text{or } \frac{1.25}{w_2} = \frac{31.8}{108} \quad \text{or } w_2 = \frac{1.25 \times 108}{31.8} = 4.25\text{g}$$

$$\text{Now } \frac{w_1}{w_3} = \frac{\text{gram equivalent of copper}}{\text{gram equivalent of iodine}}$$

$$\text{or } \frac{1.25}{w_3} = \frac{31.8}{127} \quad \text{or } w_3 = \frac{1.25 \times 127}{31.8} = 4.99\text{g}$$

2) 0.1978g of copper were deposited in 50 minutes by passing a current of 0.2 amperes through a solution of its salt. Calculate its electrochemical equivalent.

SOLUTION

We know that $Q = I \cdot t$, where Q is quantity of electricity in coulombs, I is the current in amperes and t is the time in seconds. So $Q = 0.2 \times 50 \times 60 = 600$ coulombs

600 coulombs have deposited 0.1978g of copper
hence 96,500 coulombs will deposit $\frac{0.1978 \times 96500}{600}$ g copper
 $\qquad \qquad \qquad = 31.8 \text{ g}$

So electrochemical equivalent for copper = 31.8

(Note: 96,500 coulombs of electricity are needed to liberate a gram equivalent of an element from a solution of its salt).

8. What is meant by electrode potential.
9. Explain a voltaic cell with reference to a diagram.
10. Chemical compounds which in fused state (or when dissolved in water) decompose into ions when electric current is passed through them are called :

a) Conductors b) electrolytes c) None of them

(Tick the correct one)

11. Electric conductance in metals is due to the passage of electrons through them and that in case of electrolytes it is the ions doing the same job.

12. In electrolysis, the ions liberated at the anode are called _____ and those liberated at the cathode are called _____. (Fill in).

13. In electrolysis the reduction takes place at the _____ and oxidation at the _____. (Fill in).

14. Electrolysis of aqueous solution of sodium chloride can give products other than sodium and chlorine. (correct or incorrect).

15. During electrolysis the mass of the substance deposited at the electrodes depends only on the quantity of electricity passed. (Note that it also depends on the nature of the electrolyte). (True or False).

ORGANIC CHEMISTRY

A TEACHERS GUIDE FOR SECONDARY SCHOOL LEVEL.

MAIN REFERENCE : CHEMISTRY TEXTBOOK IX-X
PAGE 245 TO 260

BY

MOHAMMAD ASLAM MALIK & KHAWAR PERVEZ MALIK
Lecturers in Chemistry, Government College of
Education for Science, Township, Lahore.

CONCEPTS TO BE DEVELOPED

1. Concept of organic compounds
2. Differences between organic and Inorganic Compounds
3. Formation of chains and rings in the organic Compounds
4. Structure of organic compounds
5. Classification of organic compounds
6. Homologous series
7. Functional groups
8. Saturated and unsaturated Organic Compounds
9. Properties of Organic Compounds
10. Concept of single, double and triple bonds
11. Organic compounds as a source of energy

OBJECTIVES

The students should be able to :

1. Identify organic compounds
2. State that organic compounds usually contain carbon and hydrogen
3. Classify organic compounds
4. Design experiment to test the behaviour of an organic compound
5. Identify carbon, hydrogen, CO_2 and H_2O .
6. Demonstrate the technique of heating organic matter
7. Recognize the effect of temperature on organic compounds
8. State that burning of hydrocarbons produces CO_2 and H_2O
9. Develop the idea of saturation and unsaturation
10. Differentiate organic compounds from each other on the basis of general properties and functional groups
11. Know the actual arrangement of different atoms in different organic molecules
12. Illustrate the difference between structural and molecular formulae.

INTRODUCTION

The teacher may give an introductory remark which includes the introduction of familiar organic compounds being used in daily life underlining the situation where chemistry is involved.

For example, at breakfast, after combing the hair, dressing clothes, wearing shoes, you and your family sit at the dining table made of wood; the chair having foam seats covered with raxine. The dining grocery set is made of plastic material, and your breakfast includes milk, bread, eggs, Jam, Cheese, sugar, tea and some fruits.

The words mentioned in the above paragraph are examples of materials containing organic compounds. These materials are largely made of carbon and hydrogen alongwith oxygen, nitrogen or sulphur, etc. such materials are called organic compounds. We can define organic compounds as those containing carbon atoms, with few exceptions like CO_2 , CO_3^{2-} , HCO_3^- , etc. because their structure and bond formation is not just like carbon compounds. Glucose, sugars, starches, oils, fats, proteins, rubber, Sui gas, vitamins, medicines are all examples of organic compounds. Originally, the term organic was used in its strict sense, since these compounds were found in living organisms, plants and animals, and could not be prepared directly. Today, however, we are able to synthesise such compounds. Urea was the first such natural organic compound to be synthesized.

TEACHERS ACTIVITY

Before going into the details of organic chemistry, the teacher should give the general concepts of organic com-

pounds through the following characteristics as compared to inorganic compounds.

1. Organic compounds catch fire easily.
2. Their melting and boiling points are generally low, with the exception of carbohydrates.
3. They are less soluble in water and more soluble in organic solvents.
4. They essentially contain carbon and, also, in most cases, contain hydrogen.
5. As a general rule organic reactions are slow.
6. They tend to have a characteristic smell and taste, e.g., esters have fruity smell and sugars are sweet in taste.
7. Many of them are essential for life process and are part of our diets.
8. Their reactivity is subject to specific conditions.
9. Most of organic compounds react by a non-ionic mechanism.
10. Molecular weight of organic compounds may reach upto thousands.

Teacher may tell the pupils that when a chemist analyses an unknown substance, he would take keen interest to know the chemical composition of the organic compound. He has a number of lines of action to adopt. He can react substances with other chemicals and see what changes occur. He can heat the substances and see what changes tend to occur on heating.

Teacher will perform simple experiments to establish the idea of heating to know the composition of the organic compounds. He may ask the pupils as to why paper on burning changes to black ash and smoke. He will explain the idea of burning and also tell about the formation of ash

and smoke, etc. He may divide the class into groups and give them different organic compounds which give ash and smoke readily, or any gas or vapour to engage the class in open discussion.

EXPERIMENT

To observe the products of burning an organic compound.

Apparatus and chemicals : Test tubes, thistle funnels, vacuum pump, bent delivery tubes, candle as an organic compound, lime water, anhydrous CuSO_4

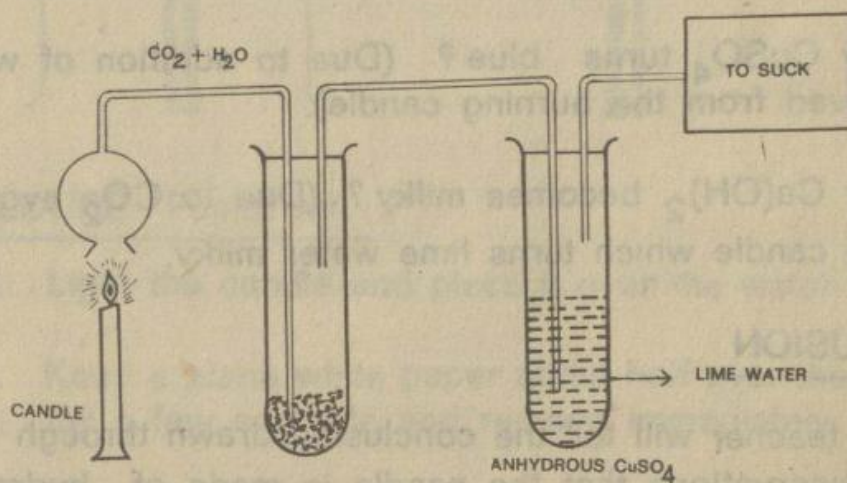


Fig. 10

PROCEDURE

- i) Wash and dry the apparatus.
- ii) Add anhydrous CuSO_4 (which has the property to absorb water) to the first tube and lime water (which has the property to turn milky when CO_2 is passed through) in other tube as shown in figure above.

- iii) Light the candle under the mouth of the inverted thistle funnel. Draw the gases through the vacuum pump.

After this experiment/demonstration the teacher will ask the pupils the following questions.

OBSERVATIONS

1. What happens to the candle ? (Candle burns, melts and reduces in size).
2. Why did thistle funnel turn black ? (Due to carbon particles produced from the candle on burning).
3. Why CuSO_4 turns blue ? (Due to addition of water evolved from the burning candle).
4. Why Ca(OH)_2 becomes milky ? (Due to CO_2 evolved from candle which turns lime water milky).

CONCLUSION

The teacher will tell the conclusion drawn through the above observations that the candle is made of hydrogen (in the form of H_2O) and Carbon.

The teacher will quote other examples from daily life to demonstrate this experiment.

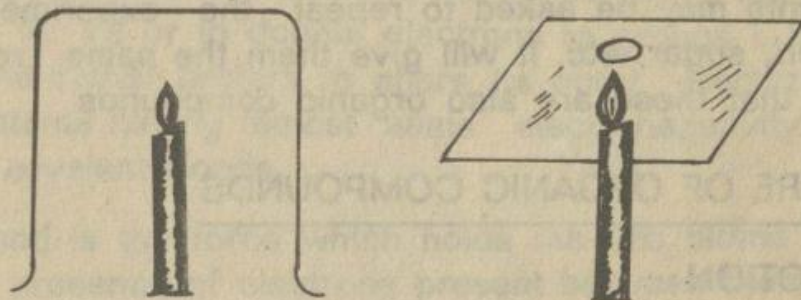
Use of sugar as food and its ultimate breakdown into CO_2 and H_2O and energy, smoking cigarette, use of fuel in vehicles, burning of match, use of cooking gas for preparing food and burning of fuel in engines are other examples.

The teacher will give simpler experiments to the pupils to perform in order to develop the understanding of what the organic compounds are made of.

STUDENTS ACTIVITY

EXPERIMENT : To study the burning of a candle.

APPARATUS & CHEMICALS : Candle, white paper, beaker and a watch glass.



PROCEDURE Pupils will :

- i. Light the candle and place it over the watch glass.
- ii. Keep a plane white paper about half over the flame for a few seconds and remove immediately.
- iii. Cover the candle with the beaker allowing the air to reach in for some time and then cover it fully with the beaker.

OBSERVATIONS

The teacher will ask the pupils as to what is the reason for the following :

- i. The candle burns and reduces in size.

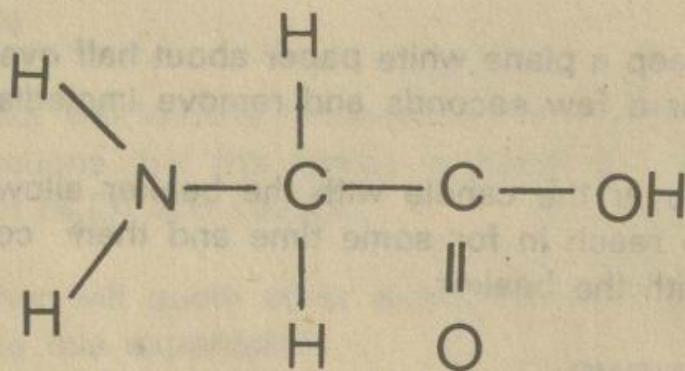
- ii. The white paper gets black. What is smoke made of?
- iii. The inverted beaker gets black.
- iv. A smoky gas is seen inside the beaker from the inner top.
- v. Vapours condense on the inner walls of the beaker.

Students may be asked to repeat the experiment by using wood, sugar, etc. It will give them the same results indicating that these are also organic compounds.

STRUCTURE OF ORGANIC COMPOUNDS

INTRODUCTION

Carbon forms a very large number of compounds due to chain mechanism reaction. The following structure is a simple two dimensional representation of a carbon compound.



BACKGROUND KNOWLEDGE FOR TEACHER

Each carbon atom forms four bonds with other atoms, nitrogen forms three bonds with other atoms, oxygen forms two bonds with other atoms and hydrogen forms single bond with other atoms.

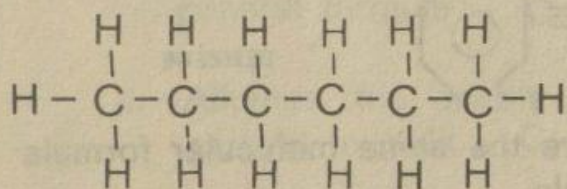
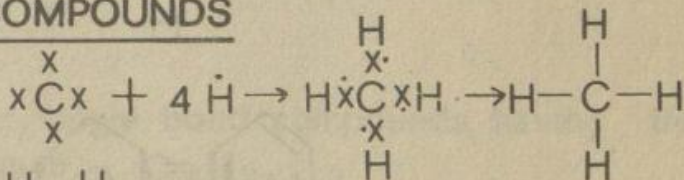
A structure like the above one gives more information about the arrangement of atoms than does the molecular formula $C_2H_5NO_2$

The structural formula also gives the concept of valency of particular atoms attached to other atoms. Valency is the combining capacity of two atoms to make a bond.

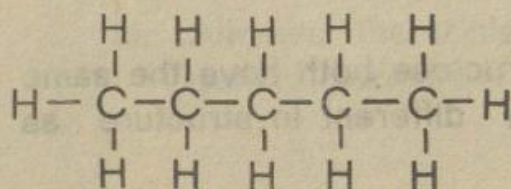
Carbon occupies a central position (group IV) in the periodic table having four electrons in its outer most shell and has no tendency, either to accept electrons from groups like V, VI, VII or to donate electrons to groups I, II, III elements. It rather prefers to share its outer electrons with other atoms having almost same electronegativity to form stable covalent bonds.

Bond is the force which holds the two atoms together by the presence of electrons present between them and is represented by a line (—). No other element can be linked to give chain compounds except carbon. That is why the number of carbon compounds is the highest. Carbon forms long chain compounds in addition to branched and cyclic compounds.

STRAIGHT CHAIN COMPOUNDS

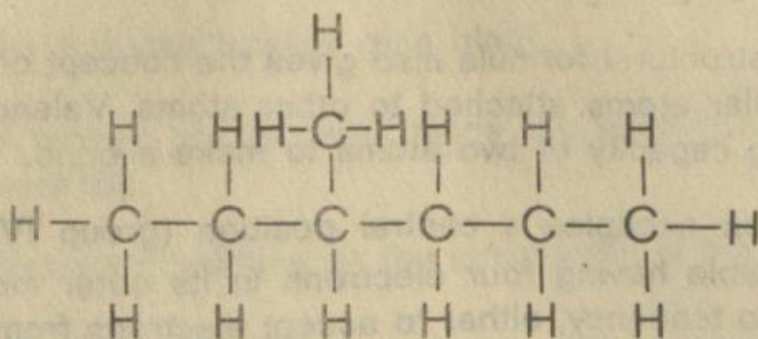


n — Hexane



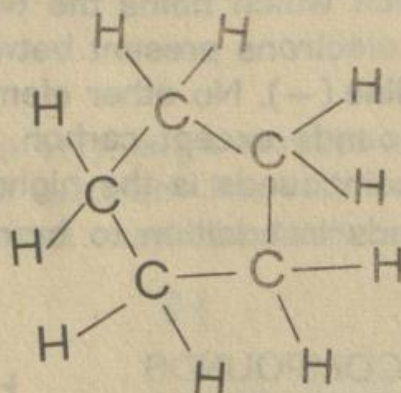
n — Pentane

BRANCHED CHAIN COMPOUNDS

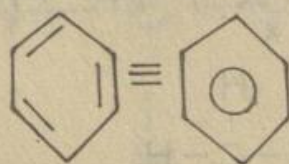


3-METHYL HEXANE

Cyclic Compounds



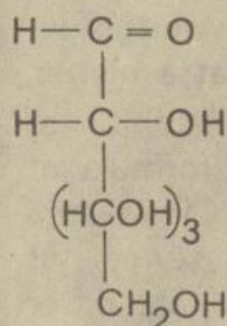
CYCLO PENTANE



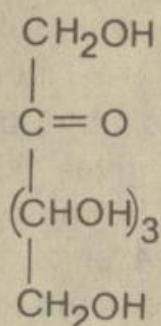
BENZENE

Organic compounds may have the same molecular formula but different structural formula.

For example, glucose and fructose both have the same molecular formula but they are different in structure as shown below :



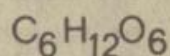
Glucose



fructose

(Structural Formula)

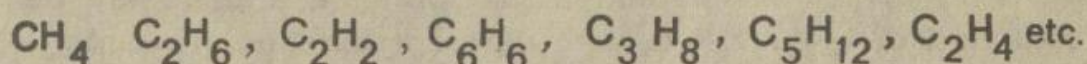
Molecular Formula
of Glucose and
fructose is



CLASSIFICATION OF ORGANIC COMPOUNDS

All organic compounds are classified into straight chain as well as closed chain compounds. These compounds contain single, double or triple bonds.

Various examples of hydrocarbons are



The hydrocarbons are divided into three families.

- i. Alkanes, the single bond compounds having the general formula : $\text{C}_n \text{H}_{2n+2}$
- ii. Alkenes, the double bond compounds having the general formula : $\text{C}_n \text{H}_{2n}$
- iii. Alkynes, the triple bond compounds having the general formula : $\text{C}_n \text{H}_{2n-2}$

While taking the simplest unit of each of the series we may make a higher unit from it. We will see that a higher compound is formed in each series by the addition of ($-\text{CH}_2-$), methylene, group. That gives the formula of successive members of the series, differing by $-\text{CH}_2-$ and thus molecular weight differing by 14 units. This series of compounds two consecutive members of which, differ by $-\text{CH}_2-$ is called **Homologous series**....

Now we come to the functional groups which give the idea about all types of organic compounds on the basis of characteristic groups present in these compounds.

Take very simple examples of different organic compounds say, R-X , R-OH , RCOOH , R-O-R , R-NH_2 , R-CHO , etc.

From the above discussion we come to understand that an organic compound is different from the other on the basis of number of carbon atoms present, and it may be saturated or unsaturated, with different functional groups.

So, on this basis they have different masses, i.e., smaller unit will have less mass as compared to larger units. Smaller units may be gases like methane, ethene or acetylene. The higher units may be liquids and still higher are solids. Hence, with the increase in size of the chain, the hydrocarbons get variation in their physical and chemical characteristics, like melting point, boiling point, etc., and in their states like gas, liquid or solid.

The parameter on which the hydrocarbons vary from each other is the functional group. Every compound belonging to the same functional group family varies in the number of carbon atoms.

The functional group for carbohydrate is (CH_2O) , which includes monosaccharides, disaccharides and polysaccharides, etc. Monosaccharides include glucose, fructose, etc., Disaccharide include sugar and Polysacchrides include starches and cellulose, etc. But we see every series differs from one another within the series due to the $-CH_2$ group, so they differ in physical states also.

TEACHER'S ACTIVITY

Teacher will define a hydrocarbon. Thus, organic compounds which are composed of only carbon and hydrogen are called hydrocarbons; for example CH_4 (methane), C_2H_6 (ethane), C_6H_6 (Benzene) etc., are all hydrocarbons.

Hydrocarbons are either saturated or unsaturated. These compounds in which all the carbon atoms in the molecule are bonded to other carbon or hydrogen atom through single bonds are called "alkanes". They are stable, less reactive and only undergo substitution reactions.

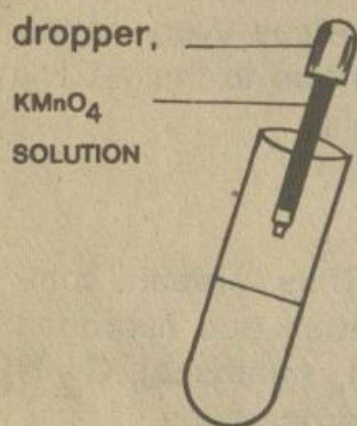
Unsaturated hydrocarbons which contain double bonds between the adjacent carbon atoms are called alkenes $(H-\underset{\substack{| \\ H}}{C}=\underset{\substack{| \\ H}}{C}-H)$ and the compounds with triple bond $(H-\underset{\substack{| \\ H}}{C}\equiv\underset{\substack{| \\ H}}{C}-H)$ are called alkynes. They have weaker π bond (alongwith the sigma bond) which can be broken. Hence, unsaturated hydrocarbon are reactive and give addition reactions on treatment with other compounds. They are said to be unsaturated because they can still add to themselves with more atoms of hydrogen or other atoms e.g. bromine till they get saturated.

TEACHER'S DEMONSTRATION

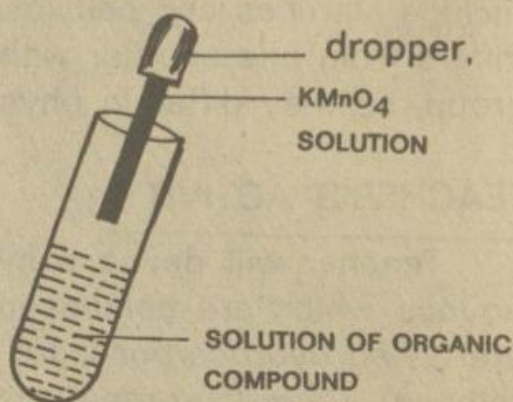
OBJECTIVE

To develop the concept of saturation and unsaturation.

Experiment Treatment of an organic compound with KMnO_4 solution.



T.T. No. 1



T.T. No. 2

Apparatus and chemicals

Test tubes, dropper, solution of Organic compounds.

PROCEDURE

- Prepare the solutions of the given saturated and unsaturated organic compounds.
- Add 2-3 ml of the saturated compound solution in the test tube 1 and add the solution of unsaturated organic compound in test tube 2.
- Add a drop of KMnO_4 Solution in each test tube and observe what happens to the colour of KMnO_4 solution.

STUDENTS ACTIVITY

Teacher will give two organic compounds to the pupils, one saturated and other unsaturated. He will ask them to

repeat the above experiment. They will be given bromine water instead of KMnO_4 solution to identify which of the compounds is saturated and which is unsaturated.

OBJECTIVE To develop understanding and practical skill in identifying the saturated and unsaturated organic compounds.

Experiment Treatment of Organic compounds with bromine water.

Apparatus and Chemicals Two test tubes, dropper, Organic compounds and bromine water.

PROCEDURE

- i. Wash and dry the test tubes.
- ii. Prepare solutions of the given compounds separately.
- iii. Add 2-3 ml of the organic liquids in the test tube 1 and test tube 2 separately.

While adding the drops of Br_2 water, the teacher should ask the pupils :

- i. What is the colour of the bromine solution ?
- ii. What happens to the colour of solution after adding in each tube ?

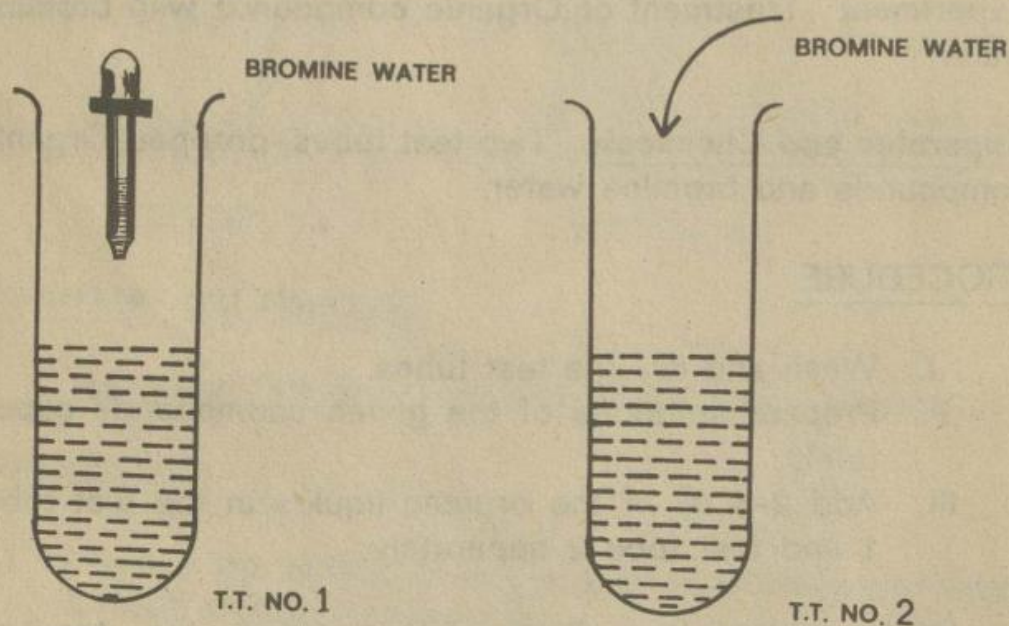
RESULTS (BASED ON OBSERVATIONS)

The teacher will tell the pupils that in test tube 1 the colour of bromine water does not discharge due to the reasons that it is a solution of saturated organic compound which does not show any addition reaction.

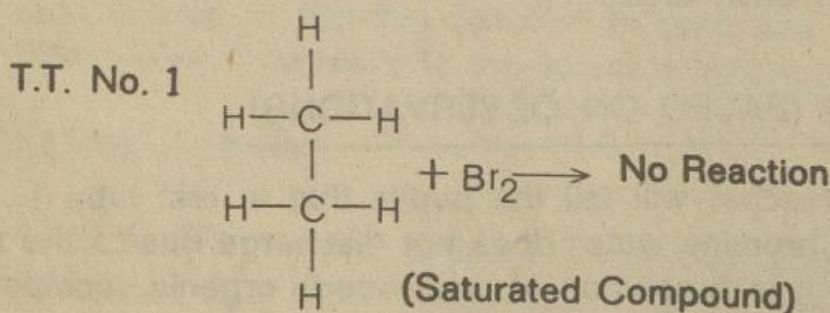
In test tube 2, colour of bromine water is discharged which indicates that the compound in this test tube is unsaturated.

CONCLUSIONS

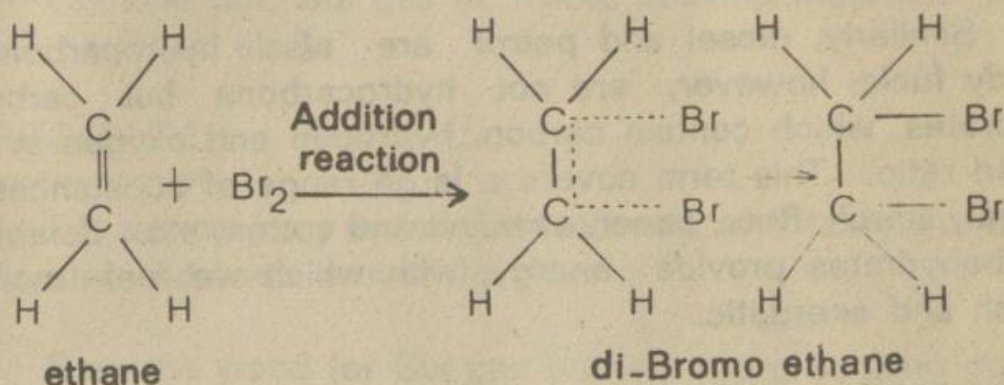
- Test tube 1 contains a saturated organic compound.
- Test tube 2 contains an unsaturated organic compound.



Teacher at this stage will tell the pupils, how and why the colour of bromine water is discharged. In unsaturated compounds, the bromine water is added up to the molecule by breaking π bond as shown in the following equation.



T.T. No. 2



(unsaturated Compound)

ORGANIC COMPOUNDS ARE SOURCES OF ENERGY

OBJECTIVES

To give the importance and uses of organic compounds.

INTRODUCTION

All organic compounds on burning provide heat or energy. The food we eat is the fuel we need. It is burnt inside our body to produce energy. The thirty tonnes of food we eat during our lifetime contains a mixture of carbohydrates, fats and proteins. Vehicles move consuming fuel, e.g. petrol, diesel, mobile oil, etc., which are converted into energy. All the above energy sources are organic materials or organic compounds.

There are many simple experiments which will show that when carbon and some of its compounds combine with oxygen energy is released. Indeed organic compounds are our principle sources of energy. We use the energy liberated from the burning fuel gas to cook our foods, and to

warm our rooms. Sui gas is a simple hydrocarbon, a combination of carbon and hydrogen only.

Similarly, diesel and petrol are also hydrocarbons. Body fuels, however, are not hydrocarbons but carbohydrates, which contain carbon, hydrogen and oxygen in a fixed ratio. This term covers a large range of substances, sugar, starch, flour, paper, sawdust and cotton, etc. Eatable carbohydrates provide energy, with which we feel more fresh and energetic.

TEACHER'S ACTIVITY

There are so many examples of the processes in our daily life which indicate that organic compounds provide energy. To make this concept clear the teacher may give the following demonstration.

EXPERIMENT: COOKING OF FOOD

- Concepts:
- i. Organic compounds are source of energy.
 - ii. The energy or heat brings changes in the other substances.
 - iii. Convert things into eatables and useables.

OBJECTIVES

1. To establish understanding of the importance of organic compounds.
2. To tell how the energy can be produced and used.
3. To learn skill techniques in burning, heating or cooking the substances.

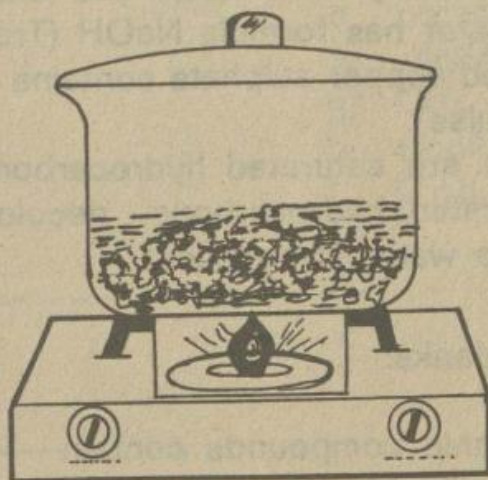
APPARATUS

Cooking pot, Sui gas or wood, cooking materials (e.g. rice, potatoes, etc.).

PROCEDURE

Put the cooking material, mixed with a suitable quantity of water and keep it over the stand.

2. Burn the wood (or Sui gas burner) under cooking pot.
3. Remove the wood or Sui gas when the material has been cooked.



Students may be asked to observe :

- i) What changes occurred during cooking ?
- ii) What has made the food cooked ?

RESULT: The wood or Sui gas being organic compounds provide energy which cooks the food and makes it eatable.

STUDENT'S ACTIVITY

Pupils may be asked to boil an egg, using the same technique.

ASSESSMENT — EVALUATION

To assess the understanding and knowledge of students, the teacher may give the objective type questions concerning organic compounds.

1. Mark as 'true' or 'false' each of the following statement:
 - a. Organic compounds contain carbon (True/false).
 - b. Organic compounds catch fire easily (True/false).
 - c. Lime water has formula NaOH (True/false).
 - d. Hydrated copper sulphate contains water molecules (True/false).
 - e. Alkenes are saturated hydrocarbons (True/false).
 - f. Unsaturated hydrocarbons decolourise KMnO_4 / Bromine water (True/false).

2. Fill in the blanks.
 - a. All organic compounds contain _____ and _____.
 - b. Carbon dioxide is a gas which turns lime water _____.
 - c. Organic compounds are source of _____.
 - d. Two examples of organic compounds are _____ and _____.
 - e. Colour of bromine water is _____.
 - f. Colour of the KMnO_4 solution is _____.
 - g. The simplest unit of hydrocarbon is _____.
 - h. Saturated hydrocarbons contain _____ bond.

(Single, double, triple)

- i. Hydrocarbons are composed of _____ and _____ only.

(C, H, O, S, N, Cl).

ANSWERS

- Q. 1 a, True
b, True
c, False
d, True
e, False
f, True

- Q.No. 2 a, C and H
b, milky
c, energy
d, sugar, Sui gas
e, organic
f, violet
g, CH_4
h, single
i, C and H
-

ABDUL SATTAR MEMON
MENGHRAJ-A-BANSARI

ATOMIC STRUCTURE AND CHEMICAL BONDING

INTRODUCTION

Man from the beginning has been trying to investigate different chemical changes and the possible explanations for these changes. The work of Dalton, Rutherford, Bohr and many other scientists has greatly increased our understanding of the nature of matter and the chemical changes occurring in matter. Our modern knowledge about the nature of matter and the chemical changes is based on our concepts about the structure of atoms and of the forces involved in binding the atoms together in molecules and bulk of matter. The objective of this unit is to illustrate how these concepts have been developed by the scientists and how they can be presented to students of secondary classes.

BACKGROUND INFORMATION FOR TEACHERS

In 1808, John Dalton, an English school teacher, started a movement of atomic theory in his famous thesis "The new system of Chemical philosophy". At that time, matter was thought to be composed of small particles. Dalton in his theory described that:

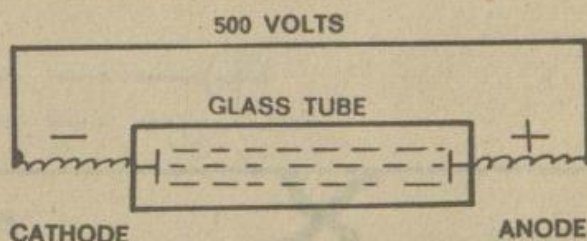
- i) Matter consists of large number of atoms. An atom is the smallest particle of an element;
- ii) The atoms of any element are identical in shape, mass and volume with respect to each other, but they differ from the atoms of other elements ;

- iii) Chemical changes take place when atoms combine together or separate from one another;
- iv) Atoms of different elements combine in a fixed proportion.

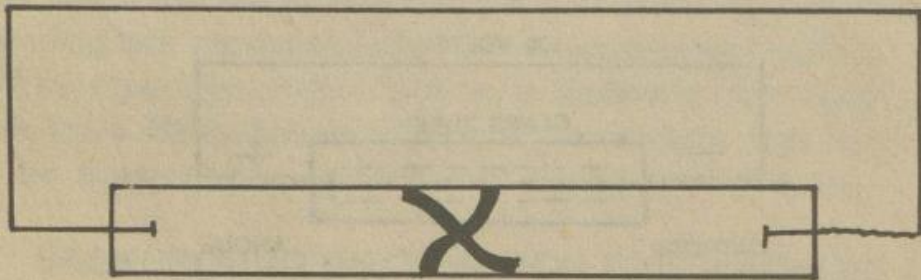
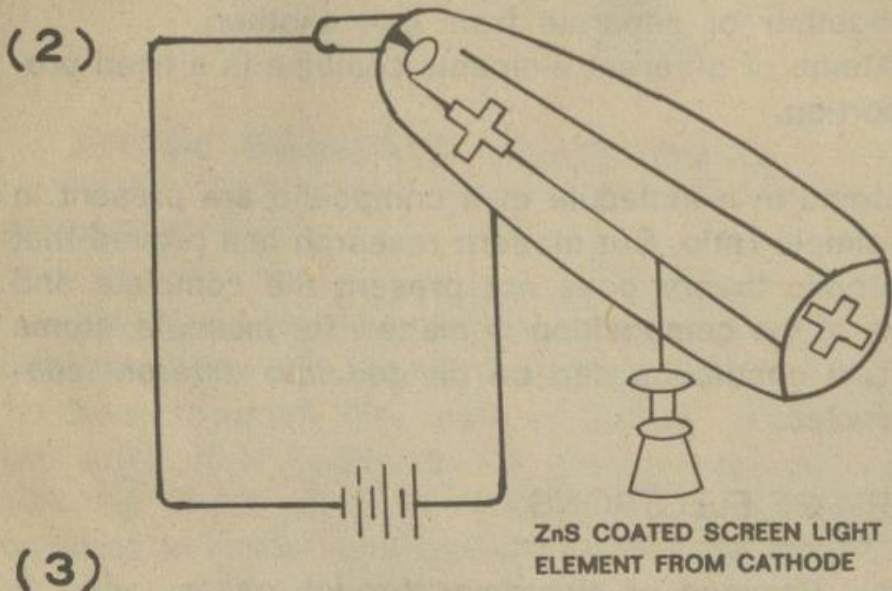
The atoms in a molecule of a compound are present in a definite simple ratio. But modern research has proved that Dalton's atomic theory does not present the complete and true picture of the composition of matter; for example, atoms under certain conditions can be divided into different sub-atomic particles.

DISCOVERY OF ELECTRONS

Experiment: Passage of electricity through gases.



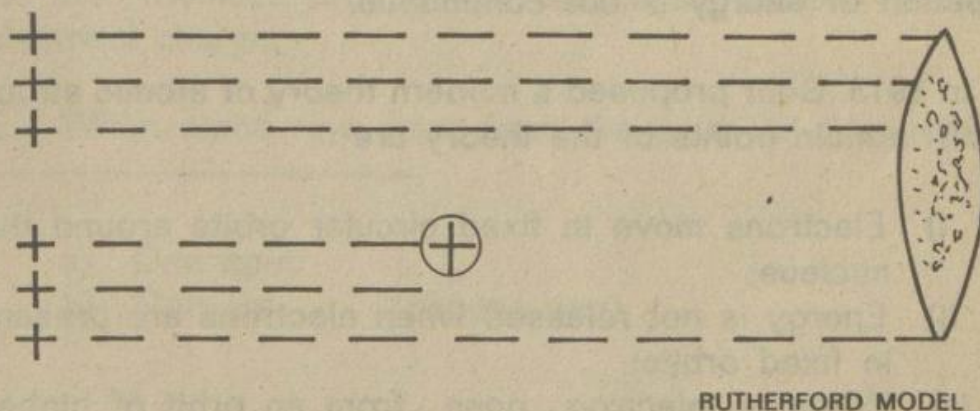
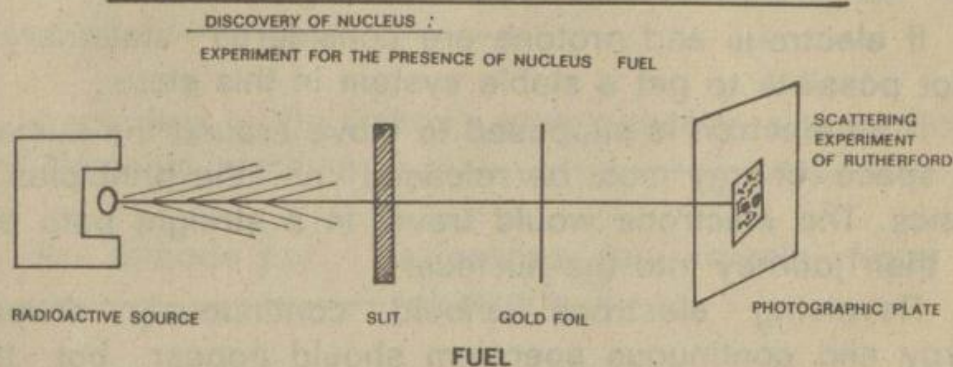
If a gas is placed in the tube shown above and its pressure is reduced, a blue glow is seen to be produced from the cathode. The negative glow was considered to arise because of 'electrons'. In 1897, J.J. Thomson found that the mass and charge of this particle was constant no matter what substance was chosen for the cathode or the gas. This means that there must be a fundamental constitution of all atoms, called electrons. The nature of the electrons can also be illustrated by the following experiments.



In series of experiments scientists showed that these rays (1) cast shadows of object in their path; (2) can make a small paddle wheel revolve; (3) are bent by a magnetic field between the poles of a magnet and an electrical field between positively charged plates. Thomson called these mysterious rays as cathode rays. As these rays could set a small paddle wheel in motion he concluded that they consisted of streams of particles which had mass. Since they are deflected by magnetic and electric fields, he concluded that they are charged. Since cathode rays are deflected away from negative plate toward positive plate, he concluded that cathode rays are made up of negatively charged particles.

DISCOVERY OF NUCLEUS

Experiment for the presence of 'Nucleus'.



i) Earnest Rutherford performed experiment with radioactive minerals. A beam of alpha particles was passed through a gold foil. He observed that the rays were not striking on one point but few of them were deflected. Thus Rutherford suggested that most of the mass of the atom was concentrated in a tiny positive nucleus at the centre of the atom surrounded by negatively charged electrons which take up most of the space. These positively charged particles in the nucleus were called protons.

ii) Moreover, since atom is electrically neutral it must contain a positive charge equal and opposite to that of electrons.

After some time Rutherford's atomic model met with three objections:

- (i) If electrons and protons are considered stationary it is not possible to get a stable system in this state.
- (ii) If an electron is supposed to move around the nucleus in a space, energy must be released on the principles of Physics. The electrons would travel in a straight path and end their journey into the nucleus.
- (iii) Revolving electrons should continuously disipate energy and continuous spectrum should appear but this disipation of energy is not continuous.

In 1913, Bohr proposed a modern theory of atomic structure. The main points of the theory are:

- i) Electrons move in fixed circular orbits around the nucleus;
- ii) Energy is not released when electrons are present in fixed orbits;
- iii) When an electron goes from an orbit of higher energy to an orbit of lower energy, energy equivalent to one quantum is released. The magnitude of quantum of energy is equal to the difference of energy between adjacent orbits, given by

$$\Delta E = E_2 - E_1 = h\nu$$

where E_2 is the energy of an electron in orbit of higher energy, E_1 is the energy of the electron in orbit of lower energy, h is a constant which is called Planks' constant (6.625×10^{-27} erg-sec) and ν is the Frequency of the electron (cycles/s).

OBJECTIVE ASSESSMENT

1. At the time of Dalton, matter was thought to be composed of _____.
2. According to the Dalton's atomic theory, atoms of different elements combine in a _____ proportion.
3. In cathode ray osilloscope, ray travels from the _____ electrode.
4. J. J. Thomson proved that electrons have _____ electrical charge.
5. When alpha rays pass from the gold foil they meet at _____.
 - a) One spot
 - b) Corners (choose one)

BACKGROUND INFORMATION FOR TEACHER

1. Elements of group 1 ionize more easily than other metal groups, thus they form stronger ionic bond with non metals.
2. Halogen gases can form ionic as well as covalent bonds (e.g. NaCl, NaI, KBr, Cl₂, F₂).
3. Hydrogen mostly forms covalent bond.
4. Ionic bonds are stronger than covalent bonds.
5. Electrolysis of ionic compounds when in liquid form can take place.
6. Electrolysis of compounds having covalent bonds does not take place.

CHEMICAL BONDING

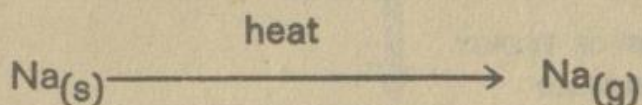
When a current is passed through molten sodium chloride in the absence of water, sodium is collected at cathode and chlorine at anode. From this simple experiment we can predict the charges on sodium and chlorine atoms when they join together in crystal of sodium chloride. The bonding between the atoms of such compounds which can electrolyse, is called ionic bonding.

In a chemical reaction, atoms always tend to acquire the electronic configuration of the nearest inert gas. When hot, sodium metal and chlorine gas are allowed to react together, it is observed that sodium reacts with chlorine and heat and light are produced.

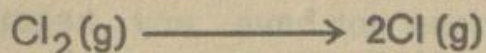


The above reaction can be explained like this :

i) Sodium (solid) and chlorine (gas) are combined and heated together; sodium first melts and then changes to a vapour. The heat required in this process is known as the heat of vaporisation.

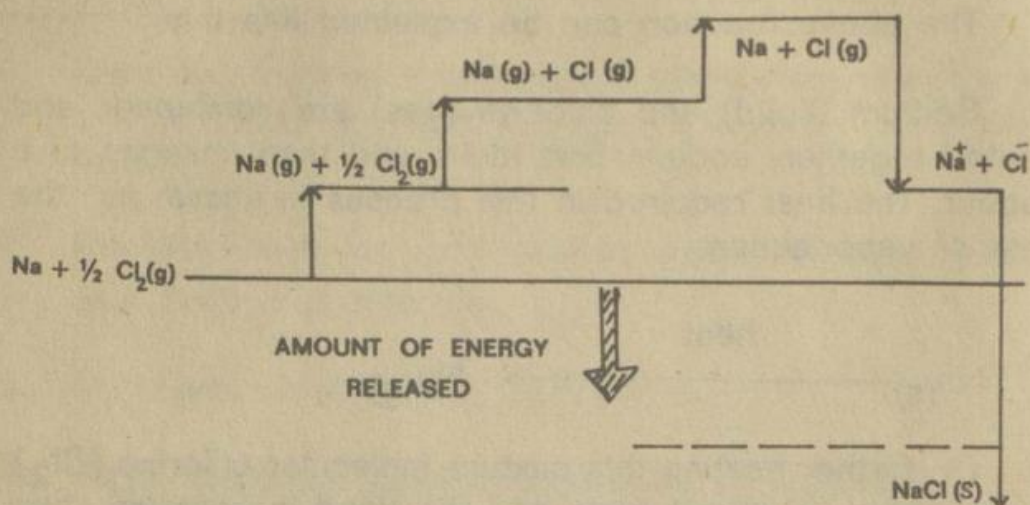


ii) On further heating this mixture molecular chlorine (Cl_2) is converted to atomic form and the bond in between two chlorine atoms ($\text{Cl}-\text{Cl}$) dissociates. The energy required in this process is called heat of dissociation.



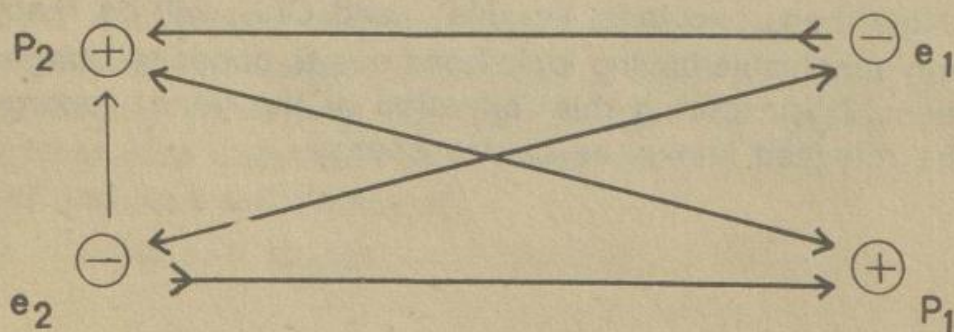
iii) After dissociation, further heat will bring the chance for outer most electron of sodium to split out and thus changing it from sodium atom to sodium ion. The energy required for this ionization process is known as ionization potential. This is the total amount of energy which has to be supplied during the reaction, but now the free electrons given out during the process come in contact with atomic chlorine which has a greater tendency to accept them. Thus during the formation of chloride ion from chlorine atom energy will be released.

Now these two ions, i.e., Na^+ and Cl^- , will be ready to form a molecule having ionic bond due to opposite charges on them. Again during this formation of the bond, energy will be released known as Lattice Energy.



When we electrolyse some other compounds (e.g. carbon tetrachloride), it has been observed that neither carbon is collected at cathode nor chlorine at anode. This simple fact shows that chlorine in CCl_4 does not have any negative charge as it was in the case of NaCl . In other words, the bond in CCl_4 is not ionic but the combination in between carbon and chlorine is of some different type. The bonding between carbon and chlorine is covalent bonding.

Let us take an example of hydrogen gas which is composed of the simplest atom, i.e. H. it has one proton and one electron. When two such atoms will come together none can give its electron because hydrogen atom has only one electron which cannot be given out as in the case of sodium. This hydrogen atom can only share the electrons and this sharing is known as covalent bonding.

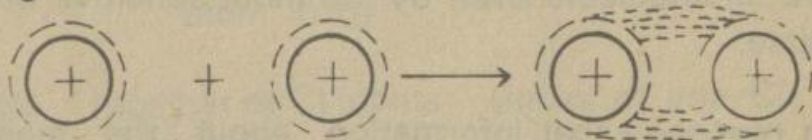


When two such hydrogen atoms come closer, the following forces are operative between them.

	<u>Attraction</u>	<u>Repulsion</u>
i) Electron-proton attraction	(i) $e_1 - p_1$	_____
ii) Electron-Electron repulsion	(ii) $e_1 - p_2$	(i) $e_1 - e_2$
iii) Proton-proton repulsion	(iii) $e_2 - p_2$	(ii) $p_1 - p_2$
	(iv) $e_2 - p_1$	

Due to greater force of attraction in between two hydrogen atoms (as shown in the figure) they will come closer to one another but we know that with the decrease of distance between them, the force of repulsion will increase. At a certain distance, where both forces will be equal, the equilibrium bond distance is called bond length.

As described in the quantum atomic theory, electrons are not found to reside in specific shells but in the form of a cloud. Similarly, after the formation of covalent bond as in hydrogen gas the electron cloud overlaps due to approaching atoms as shown in the following figure.



STUDENT ACTIVITY

1. During reactions atoms always tend to acquire the electronic configuration of _____.

2. During the formation of NaCl, sodium is heated and changes to vapour. The heat required is known as _____

- Ionization potential
- Heat of dissociation
- Heat of vaporization

3. When sodium ion and chloride ion combine together, the energy released is known as _____.
4. The bond distance within hydrogen molecule is known as _____.
5. When hydrogen molecule is formed _____ are formed in the form of clouds overlapping each other.

BACKGROUND INFORMATION FOR TEACHER

1. Different elements have different atomic masses.
2. It is impossible to weigh one atom but we instead find the relative mass of atoms.

FINDING ATOMIC MASS

Finding the atomic masses, and counting the atoms, have seemed impossible. Dalton couldn't even count them since they were undetectable even by the most sensitive balances of that time.

Dalton reasoned that information about the masses of atoms could be obtained from mass proportions of the component elements in compounds if molecular formulas were known.

Let us apply Dalton's reasoning to the compound hydrogen chloride (HCl) with mass proportions of 97.2% chlorine and 2.8% hydrogen. Assume that the compound HCl contains the simplest possible molecule with one hydrogen atom bonded to one chlorine atom H-Cl. There would then be just as many hydrogen atoms in 100g of the compound as there are chlorine atoms. Then, if a hydrogen atom had the same

mass as that of a chlorine atom, the mass proportions of the two elements in the compound would be 50% chlorine to 50% hydrogen. But these proportions were found by measurement to be 97.2% chlorine to 2.8% hydrogen, meaning that chlorine atoms don't have the same mass as hydrogen atoms. From the known mass proportions you can see that each chlorine atom must have a much greater mass than each hydrogen atom. In fact, it must have $97.2/2.8$, or about 35 times the mass.

Another way to reach the same conclusion is found algebraically. Let x be equal to the number of HCl molecules in 100g of the compound. Then x also equals the number of Cl atoms in 97.2 g of chlorine, and x equals the number of H atoms in 2.8 g of hydrogen.

$$\text{Also, } \left(\begin{array}{l} \text{number of H atoms} \\ \text{in 2.8 g of hydrogen} \end{array} \right) \left(\begin{array}{l} \text{mass of one H} \\ \text{atom in g} \end{array} \right) = 2.8\text{g}$$

Substituting: x for the number of H atoms, we get:

$$x / \left(\begin{array}{l} \text{Mass of one H} \\ \text{atom in g} \end{array} \right) = 2.8 \text{ g} \dots\dots\dots (1)$$

$$\text{Similarly, } \left(\begin{array}{l} \text{number of Cl atoms} \\ \text{in 97.2g of chlorine} \end{array} \right) \left(\begin{array}{l} \text{mass of one Cl} \\ \text{atom in g} \end{array} \right) = 97.2 \text{ g}$$

$$\text{or } x / \left(\begin{array}{l} \text{mass of one Cl} \\ \text{atom in g} \end{array} \right) = 97.2 \text{ g} \dots\dots\dots (2)$$

$$\text{Divide (2) by (1)} \quad \frac{x \left(\begin{array}{l} \text{mass of one atom of Cl} \end{array} \right)}{x \left(\begin{array}{l} \text{mass of one atom of H} \end{array} \right)} = \frac{97.29}{2.8\text{g}}$$

$$\text{or } \frac{\text{mass of one Cl atom}}{\text{mass of one H atom}} = \frac{35}{1}$$

Thus the mass of a chlorine atom is found to be about 35 times that of a hydrogen atom. Though Dalton could not determine the actual masses but he could calculate their relative masses including the lightest atom, hydrogen, set at 1, as the standard for atomic masses. Today the standard for comparison is the mass of carbon-12 atom.

A
UNIT
ON

GASES / LIQUIDS / SOLIDS

AUTHORS

1. MR. ABDUL GHAFOOR MALIK
GOVERNMENT COLLEGE OF EDUCATION
FOR MEN,
LAHORE.
2. MR. MUKHTAR AHMAD QURESHI
EDUCATION EXTENSION CENTRE
LAHORE.

LIST OF CONCEPTS

1. Three states of matter
 2. Conservation of volume by solids and liquids
 3. Compressibility of the gases
 4. Concepts of attraction between the particles in solids, liquids and gases
 5. Density of gases
 6. Pressure exerted by gases
 7. Diffusion of gases
 8. Effect of temperature
 9. Change of state of matter
 10. Pressure (relationship between force and area)
 11. Atmospheric pressure and Torcelli's Experiment
 12. Boyle's law
 13. Charle's law
 14. Absolute temperature
 15. Gas Equation
 16. S.T.P.
 17. Graham's law of diffusion
 18. Kinetic theory of gases
 19. Evaporation
 20. Vapour pressure
 21. Comparison of vapour pressure of different liquids at the same temperature
 22. Boiling point
 23. Density of solids
 24. Vapour pressure of solids
 25. Crystal lattice
 26. Structure of NaCl and CuSO_4
-

1. Concept

Particulate nature of matter.

Objectives The students will be able to STATE that matter consists of particles.

Teachers' Demonstration

The teacher can perform the following demonstration to recapitulate the concept.

- i) Diffusion of coloured crystals in water (e.g., KMnO_4 and CuSO_4)
- ii) Diffusion of perfume in the class atmosphere.
- iii) No change of volume is observed when a spoon of table salt is added to a glass full of water.

Student's Activity Fill a balloon with air and keep overnight; observe what happens and infer result from the observation.

2. Concept Change of state

Objectives: The students should be able to describe with examples how a state of matter can be changed.

Teachers' background knowledge

The students are familiar with three states of matter, they have also the experience of change of state but they have not formed the concept that most of the compounds and elements can change their state, e.g., Iron, CO_2 , etc. The teacher should remember that some of the substances under certain conditions do not behave strictly as solids, liquids or gases. Any man-made classification system will have borderline cases. The usefulness of this system of

classification far outweighs the minor complications. Some substances when heated undergo drastic changes. The production of new substances indicates a chemical change; for example, heating of sugar, when gases are cooled they can be converted to liquids, e.g., Cl_2 , O_2 , NH_3 , etc.

Teacher's Activity

1. Heating sulphur to liquid and then to vapour.

Student's Activity

Observation of change of ice to water and then boiling it to steam and drawing conclusions from the observations.

Materials Sulphur, burner, ice, test tube, china dish, a beaker. A chart showing the proximity of particles in three states of matter.

3. Concepts
 - (i) Compressibility of a gas
 - (ii) Gases exert pressure

Objectives The students should be able to give at least two examples of compressibility of a gas and the pressure exerted by a gas.

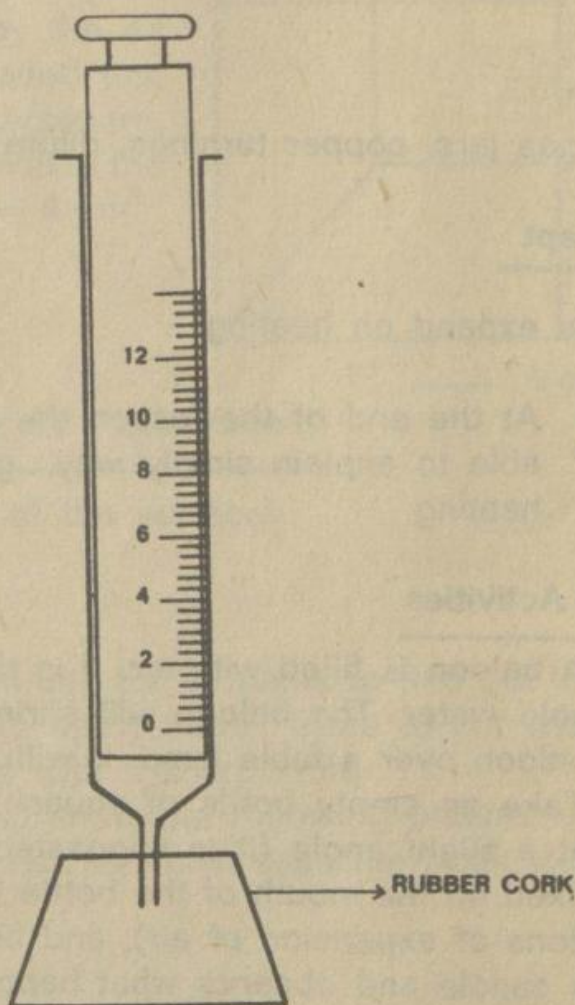
- (i) Tyre (ii) Tube (iii) Football (iv) Pump

Teacher's Activities The teacher can motivate the students by asking them about their experiences with tyres and tubes when travelling, and with pump when inflating the tube of the bicycle.

Demonstration

The springiness and compressibility of several gases can be demonstrated by drawing each gas into a plastic

syringe and compressing it with a definite force on the plunger. The degree of compression of the gas (measured by its change in volume) is observed when its springiness just balances the force on the plunger.



4. Concept

Gases diffuse freely

Objective At the end of the lesson, the student will be able to explain the process of diffusion in terms of molecular motion.

Teacher's Demonstration

Fill one cylinder with NO_2 gas and join it with other cylinder. The NO_2 will fill both cylinders. The teacher can help the students to deduce the process of diffusion.

Materials

Two gas jars, copper turnings, dilute HNO_3

5. Concept

Gases expand on heating

Objective At the end of the lesson the students shall be able to explain simply why gases expand on heating.

Student's Activities

- i) A balloon is filled with air. It is then cooled in very cold water. The balloon will shrink. On heating the balloon over a table lamp, it will expand.
- ii) Take an empty bottle of squash. Invert it in water at a slight angle (it is suggested that a straw be fixed on the mouth of the bottle for better observations of expansion of air), and heat the bottle with a candle and observe what happens.

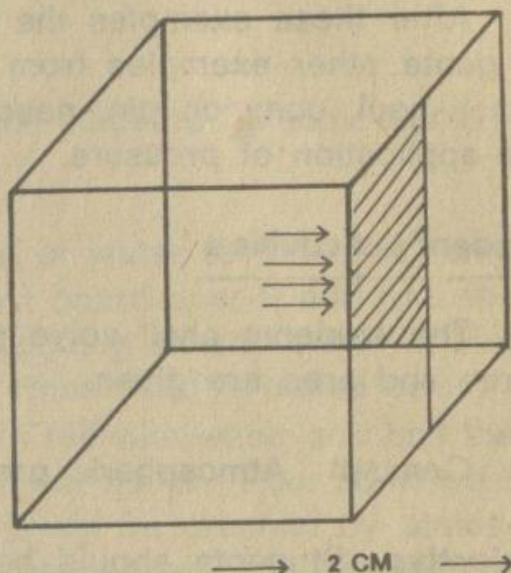
6. Concept

Pressure (as relationship between force and area) is given by $P = \frac{F}{A}$

Objectives To enable the students to calculate pressure when force and area are given.

Teacher's Note

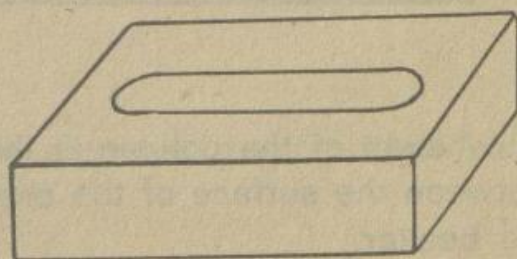
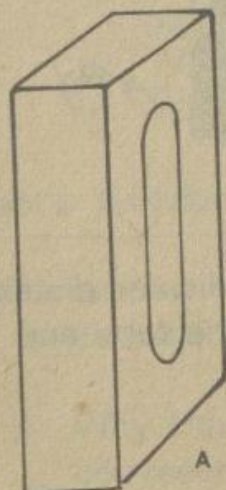
Take an empty cube of 2 Cm length with air inside; the force exerted by the air on 2 Cm² area (shaded portion) is called pressure. Total area of the side of the cube = 2cmx2cm = 4 cm²



$$\text{Pressure} = \frac{\text{Force on the surface}}{\text{Area of the surface}}$$

Discussion

- i) The teacher will ask the students to find out the reason why we have to apply more force to cut with a blunt knife compared with a sharp knife.
- ii) The teacher will show the following picture to the students and will ask in which case the brick exerts more pressure.



BRICKS

B

3. After these examples the teacher should ask students to quote other examples from their daily experiences (e.g. shoes heel, common pin, needle, -etc.) which demonstrate the application of pressure.

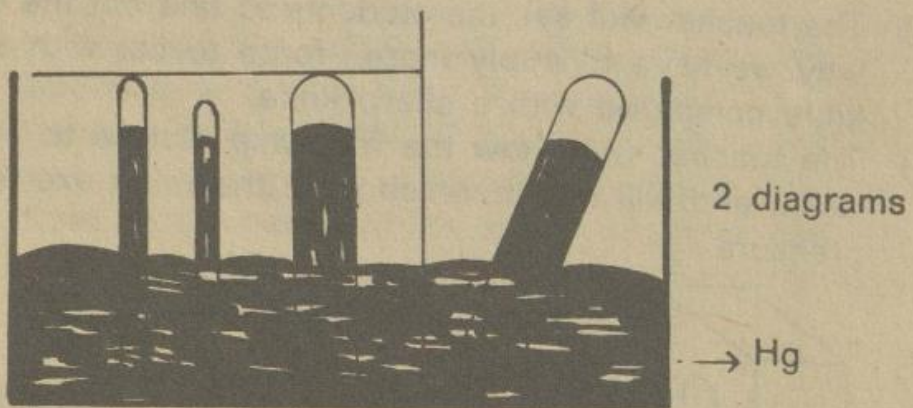
Student's Activities

The students shall solve problems on pressure when force and area are given.

7. **Concept** Atmospheric pressure

Objective Students should be able to give examples of application of the atmospheric pressure in daily life

Teacher's Note i) The diameters of the tubes in the Torricelli Experiment do not affect the height of the mercury (the pressure is working equally on each unit of area).



- ii) The height of the column is the perpendicular distance between the surface of the mercury in the tube and in the beaker.

Teacher's Activities

The teacher will motivate the students by some demonstrations as given below :

- i) Fill completely a glass of water to the brim with water. Place a thin card board over it and turn the glass upside down by putting a hand under it. The water will not run out when hand is pulled aside.
- ii) Take a tin and fill it $\frac{1}{4}$ full with water and boil the water. Place the lid to seal the tin. Pour some cold water on the hot tin, it will be crushed by atmospheric pressure. The teacher will not provide the answer to the students but will instead inquire as to what happened to the tin.
- iii) To give the "feel" of atmospheric pressure to the students, the teacher will take a plastic syringe and will call a student to pull the plunger while closing the opening with the finger or a rubber stopper. The teacher can make many variations of this demonstration and ask many probing and thought provoking questions such as :

- i) What pushes the piston back to its original position ?
- ii) Where is the force acting ?
- iii) What is responsible for this force ?

Student's Activities

The teacher will ask the students to write down the reasons for the following events :

- i) Why can we not fill the ink if there is a hole in the ink reservoir of a fountain pen ?

- ii) What pushes a soft drink up the straw and into the mouth ?
- iii) Why there is lower atmospheric pressure on a mountain top ?
- iv) Why a space suit is pressurised ?

8. **Concept:** Pressure — volume relationship for gases (Boyle's Law).

Objectives: i) The student should be able to :
 a) State Boyle's Law
 b) Solve problem based on Boyle's Law

Teacher's Note

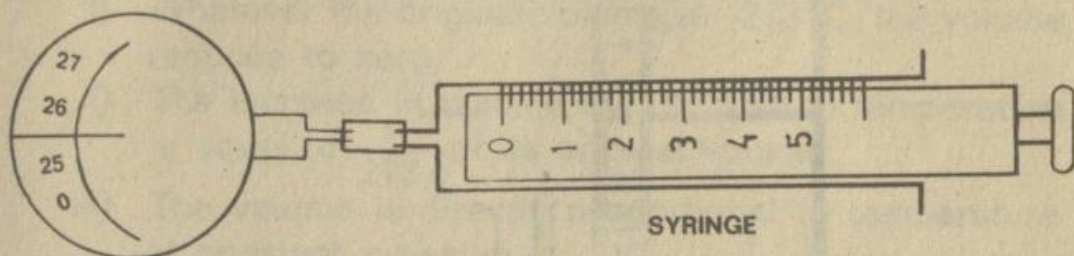
An inverse relationship between two variables means that as one increases other decreases. For example, there is an inverse relationship between the distance of two poles of a magnet and the force of attraction.

Teacher's Demonstration

The teacher can use many materials to demonstrate Boyle's Law in the class. The simplest one is that which uses a plastic syringe.

- i) Draw air into a 10 ml plastic syringe. Push in the plunger to 5 ml mark and attach the syringe tip to the dial of the pressure gauge head and record the pressure on the gauge.
- ii) Set the volume of the gas at the following readings by moving the plunger of the syringe to the following marks : 1 ml, 2 ml, 3 ml, 4 ml, 5 ml, 6 ml, 7 ml, 8 ml, 9 ml, and 10 ml. Read and record the pressure for each volume.
- iii) Use the table to discuss the relationship and ask the students to make a graph to derive an expression for Boyle's Law.

iv) Help the students to explain the phenomenon in terms of atoms and molecules.



The teacher can give an opportunity to the students to generalize from the observations, by questioning the students.

Students Activities

Solving problems based on Boyle's Law

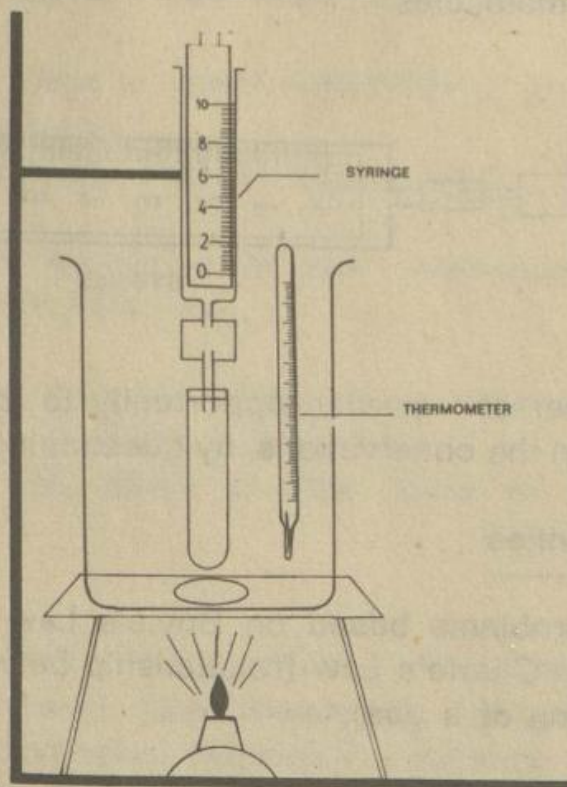
9. **Concept:** Charles's Law (relationship between temperature and volume of a gas).

Objectives

To enable the students to :

- i) State Charles's Law
- ii) Solve problems based on Charles's Law
- iii) Verification of Charles's Law with syringe and a test tube.

Arrange the apparatus as shown in the diagram. Heat the beaker and note the volume at different temperatures. Convert it to absolute scale and draw a graph to infer from it the law.

DIAGRAM

Teacher's Activity To develop the concept effectively, the teacher should discuss the following table which shows the volume of a given mass of an ideal gas at different temperatures.

Temperature (C°)	Volume (V ₁) (ml)	Volume (V ₂) (ml)	Temperature (C°)	Volume (V ₁) (ml)	Volume (V ₂) (ml)
0	273	546			
1	274	548	-1	272	544
2	275	550	-2	271	542
10	283	566	-10	263	526
100	373	746	-100	173	326
273	546	946	-273	0	0

The teacher should help the students to generalize this :

- i) Whatever the original volume, at -273°C , the volume reduces to zero.
- ii) The increase in volume by 1°C rise in temperature is equal to $\frac{1}{273}$ of its original volume.
- iii) The volume is directly proportional to temperature at constant pressure.

Students Activities

- i) To solve problems based on Charles' Law in the class.
- ii) To draw up a table similar to the above one for a mass of a gas whose volume is V at $t^{\circ}\text{C}$.

10. Concept Absolute scale

Objectives The student will be able to :

- i) Convert temperature on centigrade scale to absolute temperature scale and vice versa.
- ii) Compare the centigrade and absolute scales.

Teacher's Note

Since the volume of any mass of gas reduces to zero at -273°C , it is logical to devise another temperature scale for gases, in which zero should indicate zero volume. Hence we have taken -273°C as zero on the scale (absolute scale or Kelvin scale) and division of the scale is the same as that for the centigrade scale. The calculations involving gas laws and gas equations should be done using temperatures in absolute scale ($^{\circ}\text{K}$).

The conversion of centigrade temperature to absolute temperature is done by the equation :

$T = t + 273$ where 'T' is in absolute scale and

't' is in centigrade scale.

Teacher's Material

A chart showing both centigrade and absolute scales for comparison will be developed.

Students material activity

A graph of temperature-volume (of a given mass of gas) is given to the students to derive the relationship between temperature and volumes.

11. Concept Gas Equation

Objective: To enable the students to :

- i) Derive the gas equation
- ii) Solve problems using the gas equation

12. Concept Diffusion of gases

Objectives: The student should be able to :

- i) describe applications of Graham's Law
- ii) solve problems based on the above law

Teacher's Note

All the gases diffuse at different rates (at the same temperature). At the same temperature, the average kinetic

energy of the molecules of different gases should be the same. This is only possible when they have different velocities (the product of mass and velocity should be the same). Hence heavier molecules (CO_2) have slow velocities and lighter molecules (H_2) have fast velocities at the same temperature.

Suppose m_1 and m_2 are masses and V_1 and V_2 are average velocities of two different gases at temperature $t^\circ\text{C}$, then :

$$\frac{1}{2} m_1 v_1^2 = \frac{1}{2} m_2 v_2^2$$

$$\text{or, } \frac{m_1}{m_2} = \frac{v_2^2}{v_1^2}$$

$$\text{or, } \frac{v_2}{v_1} = \sqrt{\frac{m_1}{m_2}}$$

Since velocity and rate of diffusion, and mass and density of the gas are related together, the relative rate of diffusion of two gases is given by :

$$\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$$

This is Graham's Law

For oxygen and hydrogen, we can write :

$$\frac{r_{\text{Oxy}}}{r_{\text{Hyd}}} = \sqrt{\frac{1}{16}} = \frac{1}{4}$$

which is the ratio of their molecular weights. Hence,

H_2 diffuses 4 times faster than O_2

Teacher's Demonstration

Take a meter long glass tube and make two cotton plugs, soak one in concentrated HCl and other in concentrated

NH_4OH and plug them onto the both ends of the tube at the same time. (keep the tube stationary).

Dense fumes of NH_4Cl will quickly form. A pupil can measure the distances moved by the fume and the relative rates of movement of HCl and NH_3 molecules may be calculated.

$$\frac{r(\text{NH}_3)}{r(\text{HCl})} = \frac{72 \text{ cm}}{28 \text{ cm}} = \frac{5}{2} \quad (\text{say})$$

SAFETY Concentrated HCl and concentrated NH_4OH must not be allowed to come into contact with human skin.

Students Activities

- i) The students will fill a balloon with hydrogen and the other with air. Observe which loses the gas at a faster rate.
- ii) Solve problems based on Graham's Law.

13. Concept Kinetic Molecular Model of a gas

Objectives i) The students should be able to state the molecular model in their own words.

ii) The students should also explain, how the molecular model explains, how the motion of the molecules is increased as the temperature is increased.

Teacher's Note The kinetic picture has been derived from the observation of the motion of the dust particles. This has been very effective in explaining the behaviour of the gases. All the gas laws can be derived from the model. This model has also been used to describe liquids and solids and the effect of the temperature on them.

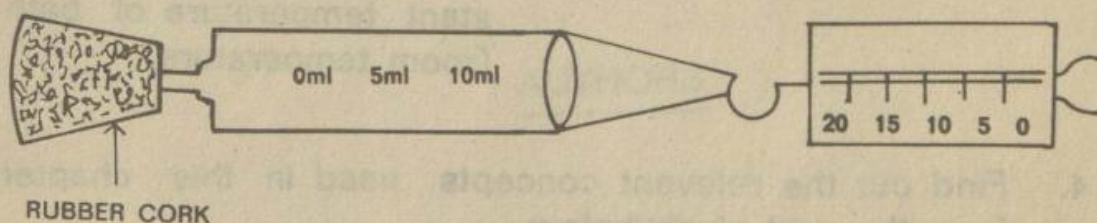
Teacher's Material

The teacher can use the mechanical model to explain the continuous and random motion of the molecules. A simple mechanical model can be a small tray or petridish and a few plastic, glass or steel beads. Overhead projector or Epidioscope can be used to show their motion on the wall.

Student's Activities Let the students make their own imaginative statement to describe the behaviour of the molecules.

Suggested Activities

- i) Measuring the Atmospheric pressure with a syringe and spring balance.

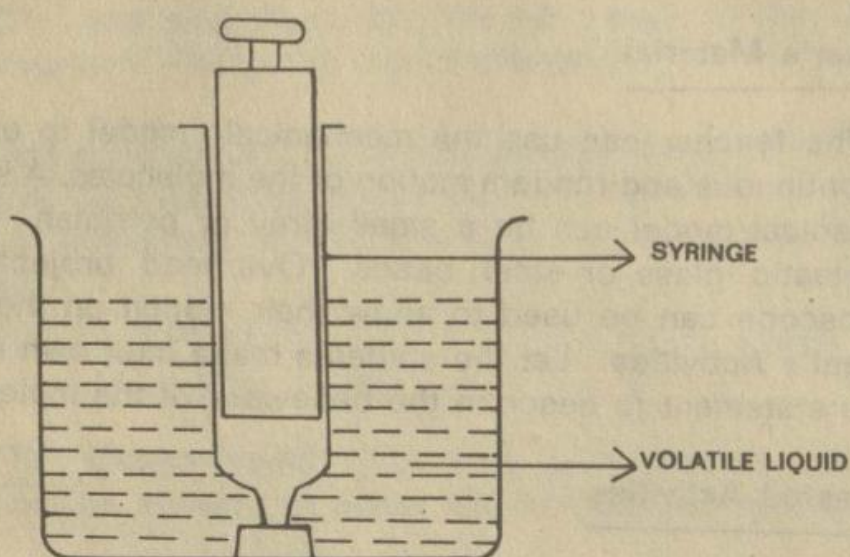


Materials Rubber cork, plastic syringe, thread and spring balance.

- ii) Change of state from liquid to gas by changing the pressure and not the temperature.

Materials: Beaker, Syringe, Rubber cork and a volatile liquid (ether).

The syringe is filled with a small amount of ether kept in the bath at room temperature and then the plunger is pulled out.



Syringe filled with a small amount of ether at constant temperature of bath (room temperature).

4. Find out the relevant concepts used in this chapter from the grid given below :

X	G	F	D	M	D	N
G	A	S	O	L	I	D
B	O	C	A	I	F	M
O	M	H	U	Q	F	O
Y	P	A	C	U	U	M
L	S	R	Y	I	S	S
L	T	L	W	D	E	T
L	U	E	T	S	O	P

TOPIC

REACTION RATES

AUTHORS

1. RAJA HAQ DAD TARIQ
LECTURER
GOVERNMENT COLLEGE OF EDUCATION
AFZALPUR, MIRPUR (A.K.)
2. MOHAMMAD YOUNES QURESHI
SUBJECT SPECIALIST
EDUCATION EXTENSION CENTRE
MUZAFFARABAD (AZAD KASHMIR)

REACTION RATES

1. INTRODUCTION

The rates at which chemical reactions occur are just as important to you as they are to industrialists, chemical engineers and agriculturists. You might be interested to know how long you keep the eggs fresh at room temperature, how can you save vegetables and fruits from spoiling during storage over long periods of time. An engineer may like to know the rate at which concrete sets, and an agriculturist may be interested in the rate of growth of a particular vegetable crop. All these problems are related to the field of chemical kinetics for the study of the rates of chemical reactions. In this unit we will study some factors which affect the rates of chemical reactions.

2. LIST OF TOPICS

1. Defining the rate of a reaction.
2. Slow and fast reactions.
3. Factors which influence the rate of a chemical reaction.
4. Collision theory of reaction rates.

3. OBJECTIVES

After completing this unit students will be able to :

- i) Describe rate of reaction in quantitative terms
- ii) Investigate the effect of concentration on the rate of reaction
- iii) Find the effect of temperature on the rate of reaction

- iv) Find the role of catalyst on reaction rate
- v) Estimate the effect of pressure on reaction rate
- vi) Find the effect of size of particle on rate of reaction
- vii) Identify the role of light on certain reaction rates
- viii) Comprehend that reactions occur because of molecular collisions and that in order to be effective the collisions must possess appropriate energy, and must have proper orientation in space.

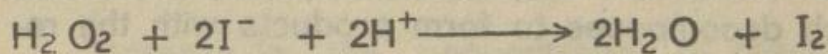
4. BACKGROUND INFORMATION FOR THE TEACHER

THE CONCEPT OF REACTION RATE

During a chemical reaction, reactants are being converted to products. The reaction rate tells us how fast the reaction is taking place in terms of the consumption of reactants or the accumulation of products in a given time. Hence,

$$\text{Reaction rate, } r = \frac{\text{Change in amount (or concentration) of a substance}}{\text{time taken}}$$

Thus we can define reaction rate as the rate of change of concentration of a particular reactant or product. For example, when acidified hydrogen peroxide is added to a solution of potassium iodide, iodine is formed :



If the concentration of iodine rises from 0 to 10^{-5} M in 10 seconds, the reaction rate is given by :

$$r = \frac{\text{Change in concentration}}{\text{time taken}} = \frac{10^{-5}}{10} = 10^{-6} \text{ Ms}^{-1}$$

Using the symbol Δ to represent the change in a particular quantity, we can write : $\frac{\Delta[I_2]}{\Delta t} = 10^{-6} \text{ M S}^{-1}$

EFFECT OF TEMPERATURE ON REACTION RATE

Magnesium, carbon, sulphur and hydrogen do not react with oxygen at ordinary temperatures. It is only at higher temperatures that these elements react. On lighting a candle, the candle first burns with a low flame, but soon it begins to burn with an appreciably high flame as the temperature rises.

At a fixed temperature, a chemical reaction rate depends upon number of fruitful collisions taking place in unit volume and in a unit time between reacting molecules. Thus at a given temperature and pressure, the number of collisions per unit time between the reacting molecules would depend on their concentration. The rate of reaction would increase with increase in concentration of reacting molecules.

EFFECT OF CATALYST ON RATE OF REACTION

A catalyst is a substance which, when used in very small amounts, increases or decreases the rate of a reaction. A catalyst itself does not change at the end of a reaction. The role of catalyst is not passive; it is involved in the formation of an intermediate species called the 'activated complex' which decomposes to form products with the regeneration of the catalyst.

EFFECT OF PRESSURE ON REACTION RATE

Chemical reactions in gaseous state are greatly affected by pressure changes. Whenever a chemical reaction in the

gaseous state is in equilibrium state, an increase in pressure will shift the equilibrium in the direction of decrease in volume. However no such change is observed in reactions taking place in liquid phases. In these cases, an increase or decrease of pressure will have little or no effect on reaction rate.

EFFECT OF SIZE OF PARTICLE ON REACTION RATE

As chemical reaction occurs due to collision of molecules, the greater the surface area of a reactant, the greater will be the number of collisions. A decrease in particle size would result in over all increase in the available surface area at which reaction can occur. Thus rate of reaction will increase by decreasing the size of particle.

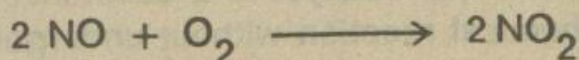
EFFECT OF LIGHT ON CERTAIN REACTION RATES

There are certain reactions which occur in sunlight. The rates of such reactions in the absence of sunlight are very slow. For such reactions the sunlight provides the requisite energy for the completion of the reaction. Many reactions proceed to completion under the effect of radiation at different wave lengths.

KINETIC EXAMPLES

The Concentration of reactants

Increase in the concentration of reactants normally causes an increase in the rate of a reaction; but this is not always the case. Furthermore, different reactants can affect the rate of a particular reaction in different ways. For example, consider the reaction between nitrogen oxide and oxygen.



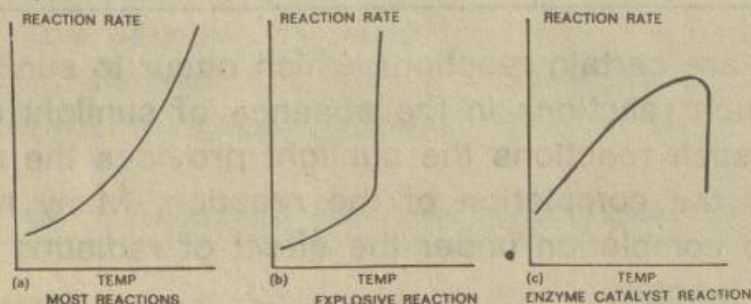
The reaction rate doubles when the oxygen concentration is doubled, but doubling the concentration of nitrogen oxide quadruples the rate of reaction.

The temperature of the reactants

Milk and other perishable foods spoil much more rapidly in summer than in winter. In summer the chemical reactions in the decomposition process occur more rapidly as the temperatures are quite high. In general, increasing the temperature increases the rate of chemical reactions.

Reactions speed up at higher temperature.

Look at the graphs of reaction rates against temperature in the figures given below and answer the questions that follow.



1. Why is there a sudden change in the rate of an explosive reaction at one particular temperature.
2. The rate of an enzyme catalysed reaction increases at first, but decreases as the temperature increases. Why is this so?

Catalyst

As noted before, a catalyst is a substance which alters the rate of a chemical reaction without undergoing any over-

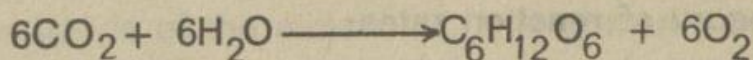
all chemical change itself. Although catalysts are not consumed in reactions they are believed to participate by forming intermediate compounds in conversion of reactants to products.

Normally catalysts are used to accelerate the reactions. Certain catalysts can, however, be used to slow down reactions. For example 1, 2, 3 Propane-triol (Glycerine) is sometimes added to hydrogen peroxide as a negative catalyst in order to slow down its rate of decomposition. Catalysts play an important part in biological and industrial processes by enabling reactions to take place which would never occur in their absence. Many industrial processes including the manufacture of ammonia, sulphuric acid, nitric acid, ethane, polythene and polystyrene rely on the use of a catalyst. Living things rely even more heavily on catalysts because every chemical reaction in animals, plants, and micro-organisms requires its own specific enzyme acting as the catalyst.

LIGHT

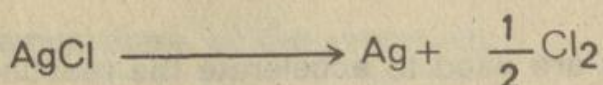
Photosynthesis and photography both involve light-sensitive reactions. The leaves of plants contain chlorophyll, a green pigment, which can absorb radiation in the visible region of the electromagnetic spectrum and use this energy to synthesize chemicals and provide food for the plant.

During photosynthesis, plants transfer carbon dioxide and water into oxygen and sugars (such as glucose).



In the absence of sunlight, energy is no longer provided and photosynthesis ceases.

White AgCl turns purple and finally darkgrey when it is exposed to sunlight which provides the energy required to decompose the silver chloride.



The use of silver salts in photography depends on photo-sensitivity of this kind. The reactions of halogens with hydrogen and with alkanes are other examples of photochemical reactions.

REACTION RATES AS A FUNCTION OF TEMPERATURE AND CONCENTRATION

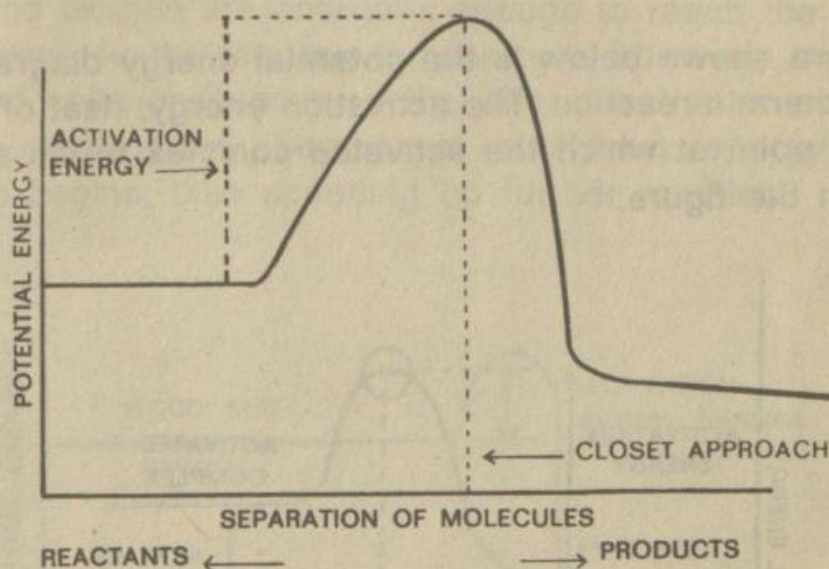
To explain the temperature effect on reaction rates, let us use the kinetic molecular theory. It seems reasonable to assume that molecules react only when they collide. You have learned that raising the temperature of a group of molecules increases their average kinetic energy. This results in more collisions between molecules in a given time, and the collisions are harder and effective on the average. Both factors would tend to speed up the rate at which colliding molecules react.

Similarly, if the concentration of reactant molecules is increased, there is a greater number of collisions per second taking place between molecules, and they are more likely to react. This explains the effect of concentration of reactants on reaction rates. Such explanations are derived from the collision theory of reaction rates.

ACTIVATED COMPLEX AND ACTIVATION ENERGY

To gain a better understanding of molecular reactions taking place on account of collisions, it is believed that there

must be a minimum potential energy increase during collision before a reaction can occur between two molecules. This minimum increase in potential energy is called the activation energy.



"Potential energy changes in a typical exothermic reaction."

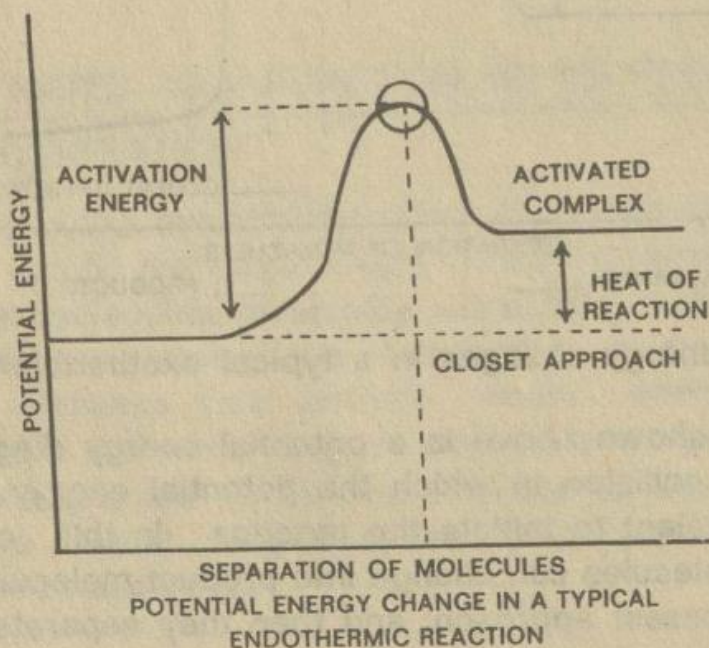
The figure shown above is a potential energy diagram representing collision in which the potential energy increase is just sufficient to initiate the reaction. In this case the reactant molecules can change into product molecules at the point of closest approach, and then they separate.

Notice in the figure that the potential energy sum of the separating product molecules is different (in this case less) from the potential energy sum of the approaching reactant molecules. This represents the difference in stored energy in the bonds of the molecules. It is primarily this difference that accounts for the heat of reaction of a reacting system. The reaction represented in figure is exothermic.

The group of colliding molecules at the instant of changeover from reactants to products is called an activated

complex. The activated complex is unstable because of its high potential energy. It, therefore, exists only fleetingly at the instant of reaction. The activated complex for any given reaction step is believed to have a definite geometric shape.

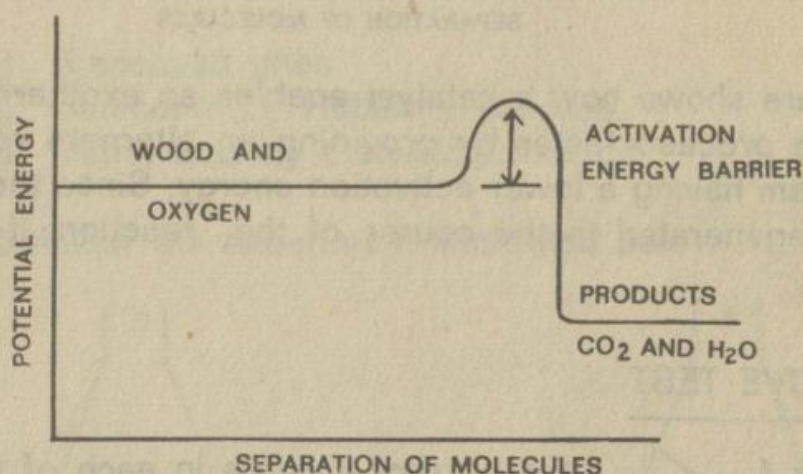
Figure shown below is the potential energy diagram for an endothermic reaction. The activation energy, heat of reaction and point at which the activated complex forms are all shown in the figure.



Let us now use these ideas to try to answer a question which may be now have occurred to you. Wood stores a lot of potential energy. Since wood is already in contact with oxygen in the air, then why it doesn't react spontaneously with the air and burn up? After-all; a rock at a high potential energy will fall spontaneously to a lower potential energy. Try to answer this question before reading further on.

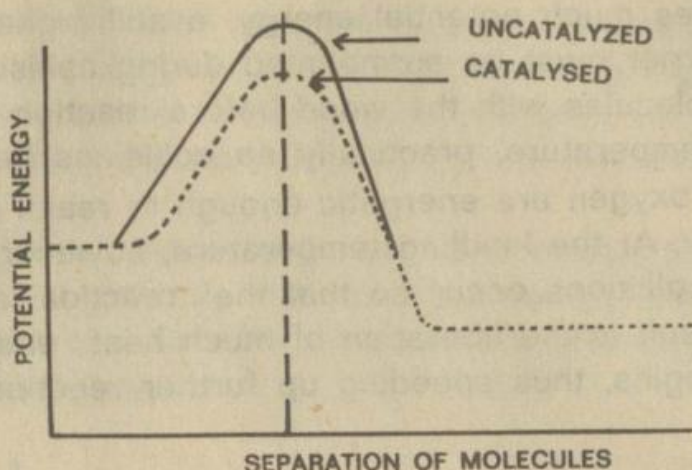
According to the theory given above, the activation energy acts as a sort of barrier to a reaction. Even though

wood stores much potential energy, a still higher potential energy barrier must be surmounted during collision of the oxygen molecules with the wood before reaction can occur. At room temperature, practically no collisions between the wood and oxygen are energetic enough to reach the activation energy. At the kindling temperature, however, sufficient activated collisions occur so that the reaction is possible. The net result is the liberation of much heat energy once reaction begins, thus speeding up further reaction.



CATALYSTS

How can a substance which is not used up in a reaction cause the reaction to occur at a faster rate? This question puzzled chemists for sometime. A catalyst is believed to provide an alternative reaction mechanism with a lower activation energy. The catalyst is thought to react in one of the early steps of the mechanism and to be regenerated in a later step. Figure shown below compares the potential energy diagram of a catalysed reaction with one for the same reaction uncatalysed. The lower activation energy for the catalysed reaction results in a greater proportion of effective collisions, thus speeding the reaction rate at a given temperature. Regeneration of the catalyst prevents it from being used up as reactants normally are.



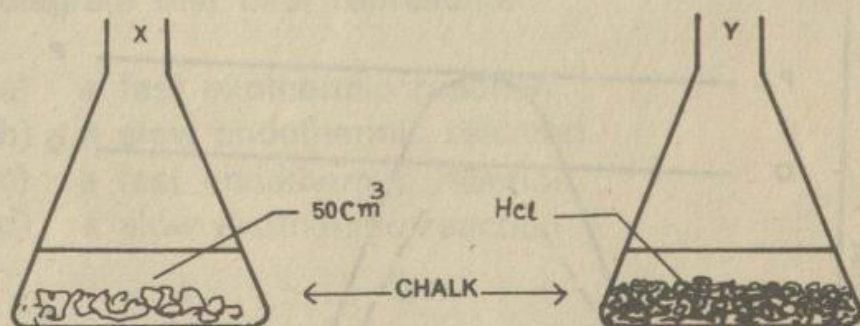
This figure shows how a catalyst enables an exothermic reaction to proceed faster by providing an alternate reaction mechanism having a lower activation energy. Since the catalyst is regenerated in the course of the reaction, it is not used up.

OBJECTIVE TEST

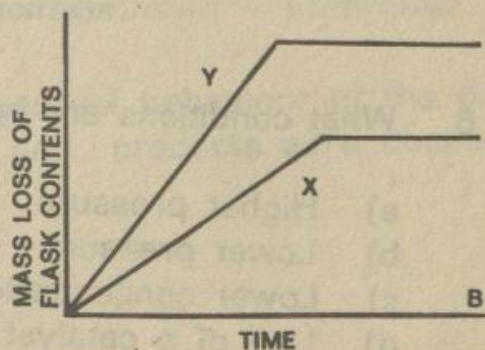
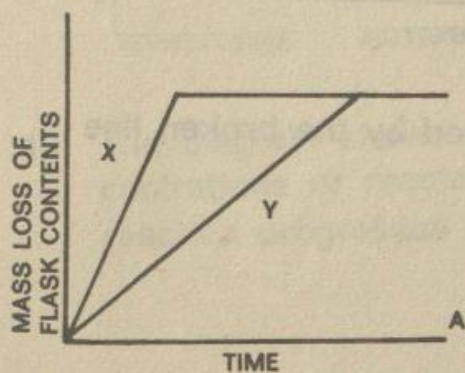
Tick (☒) the correct choices in each of the following statements :

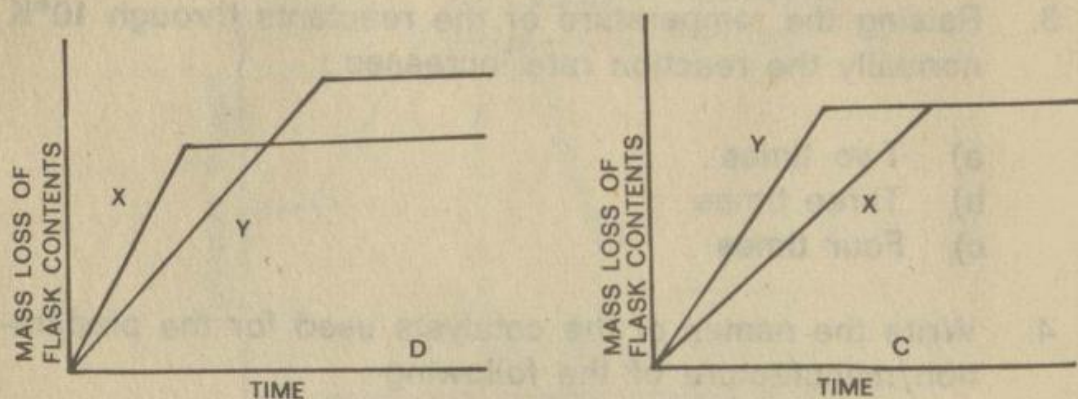
1. The rate of reaction decreases
 - a) With increase in concentration of reactants.
 - b) With decrease in concentration of reactants.
 - c) Without any change in concentration of reactants.
2. A catalyst helps a chemical reaction
 - a) Because it stops back reaction from proceeding.
 - b) It lowers the energy of the products.
 - c) It lowers the activation energy for the reaction.
 - d) It alters the equilibrium constant for the reaction.

3. Raising the temperature of the reactants through 10°K normally the reaction rate increases :
- Two times
 - Three times
 - Four times
4. Write the names of the catalysts used for the preparation/manufacture of the following :
- Vanaspati ghee
 - Ammonia by Habber's process
 - Nitric acid by Ostwald's method
5. Consider the experiment illustrated below

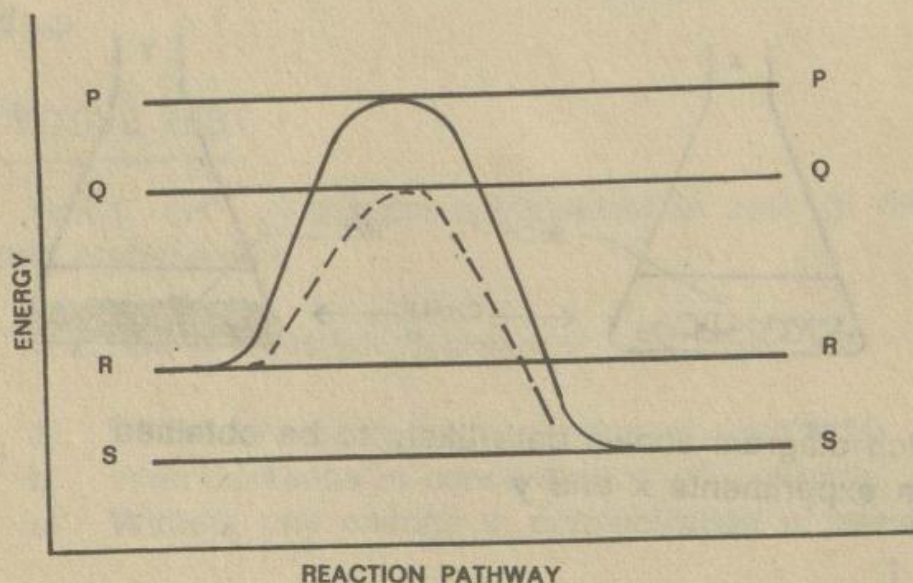


Which diagram shows data likely to be obtained from experiments x and y





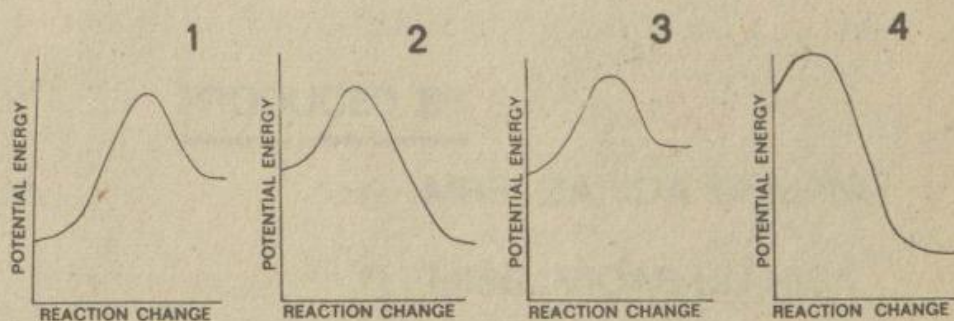
Questions 6 and 7 refer to the following diagram in which the unbroken line represents the reaction $2\text{SO}_2 + \text{O}_2 \longrightarrow 2\text{SO}_3$ and the broken line the same reaction under different conditions



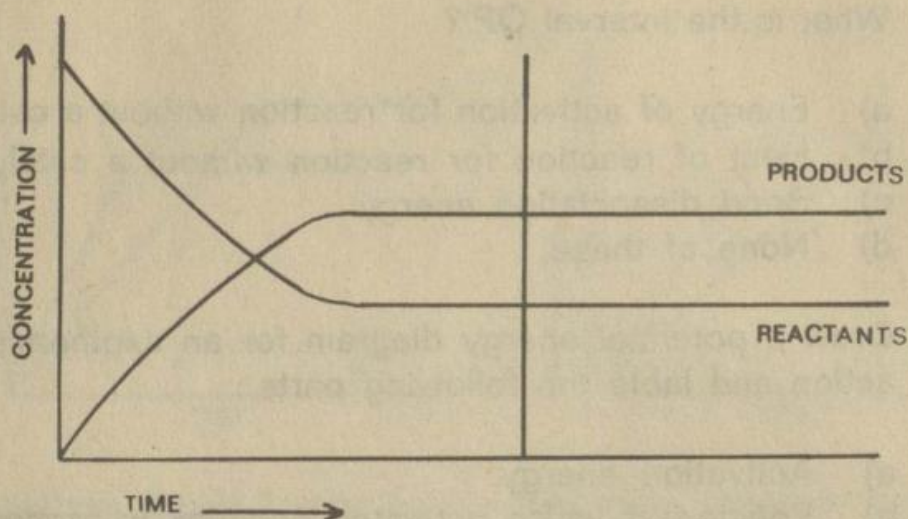
6. What conditions are indicated by the broken line

- Higher pressure
- Lower pressure
- Lower concentration
- Use of a catalyst

7. What is the interval QP ?
- Energy of activation for reaction without a catalyst.
 - Heat of reaction for reaction without a catalyst.
 - Bond dissociation energy.
 - None of these.
8. Draw a potential energy diagram for an exothermic reaction and label the following parts :
- Activation energy.
 - Position at which activated complex is formed.
 - Heat of reaction.
9. Select from the given diagrams one potential energy diagram that best represents :
- a fast exothermic reaction
 - a slow endothermic reaction
 - a fast endothermic reaction
 - a slow exothermic reaction



10. The diagram shows the general behaviour of the concentrations of reactants and products as a chemical reaction progresses :



- At what stage are the concentrations varying most rapidly?
- At what stage is the reaction at time "x"?
- What happens to the rate of change of concentration of the reactants as the reaction proceeds?

A

MODEL UNIT

ON

"M O L E"

PRODUCED BY

1) MRS. ZAHIDA USMANI

2) MISS. VIQAR-UN-NISA

MOLE

INTRODUCTION

By introducing the mole, the students should gain an understanding of the small size of atoms, formulae, relative atomic mass, ratio and weights of different atoms.

As we have specific units in our daily life, e.g. one pair of socks, one dozen of eggs, etc., we have in a similar way, a unit for counting atoms, ions, molecules and electrons. This unit is called the "mole" and it indicates a definite quantity of particles. It indicates the number of atoms in one gram atom or the number of molecules in one gram molecule.

CONCEPTS

1. Definitions
2. Ratio, relative atomic mass and size of atoms
3. Avogadro's constant
4. Exercise
5. Electro-chemistry
6. Solution
7. Molar volume
8. Exercises pertaining to molar volume
9. List of substances having formula masses as whole number
10. Exercises in connection with the substances having formula masses as whole numbers
11. Faraday's Laws

SPECIFIC OBJECTIVES

At the end of this topic the pupil should be able to :

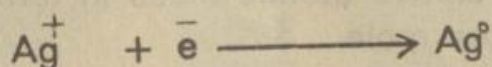
1. Define relative atomic mass in terms of carbon-12.
2. Define the mole
3. Define molar mass
4. Calculate the molar mass of a substance, and relative atomic masses of the elements
5. Calculate the amount (number of moles) of a substance, given its mass and molar mass
6. Calculate the number of particles in a given amount of a substance
7. State that 1 mole of any gas at S.T.P. occupies 22.4 l
8. Calculate the concentration of a solution, given its volume and number of moles of solute
9. Calculate the amount of solute in a substance, given its volume and concentration
10. State that the number of moles of electrons required to deposit one mole of a metal from a solution of its salt is equal to the valency of that metal.

DEFINITIONS

1. Atomic weight in grams is called gram atomic weight, molecular weight in grams is called gram molecular weight and formula weight in grams is called gram formula weight. Each of these expressions in grams is called a gram Mole (g.mol).

2. The mole is defined as the amount of substance that contains as many elementary particles as there are atoms in 12g of carbon-12.

One mole of matter is liberated by passing a current of 1 Faraday or 96,490 coulombs through an electrolytic solution. For example,



Thus, the reduction of 108g (1 mole) of Ag^+ to Ag^0 requires 1 mole of electrons.

A mole of a substance contains 6.02×10^{23} particles, i.e., atoms, molecules, or ions. This number is represented by N in Chemistry and is called Avogadro's number.

3. Ratio, Relative atomic mass and sizes of atoms.

The sizes of atoms can be judged knowing that 1 mole of a substance contains 6.02×10^{23} atoms; e.g. 18g of H_2O (equal to one mole) contain 6.02×10^{23} molecules of water, assuming its density to be 1.

The concept of relative atomic mass and the concept that 1 g atom of any element contains the same number of atoms can be explained by the following examples.

A four annas coin, eight annas coin and one rupee coin are in the mass ratio 1:2:4. This can be called their relative masses.

At this point we can compare this example with the atoms of elements, introducing ^{12}C scale of relative atomic masses and hence introducing the g-atom simply as the relative atomic mass in grams.

Consider now the example of coins :

Four annas	eight annas	one rupee
1.77	3.58	7.03
1g	2g	4g

Weigh three opaque bottles (so that pupils will not know the number of coins in each bottle). Put 12 of each type of coins in separate bottles. Each of the bottles is weighed and the weights of the contents are found by subtraction of the first weight from the last recorded weight.

Four annas	eight annas	one rupee
21.46	42.84	85.09
1	2	4

Empty out the contents of the bottles to show the pupils they contain the same number of coins. Hence we can say that '1g atom' of 4 annas coins will contain the same number of atoms as 1g atom of eight annas coin.

This analogy can help to show that one g atom of any element will contain the same number of atoms as 1g atoms of any other element. Scientists call this quantity as Avogadro's constant.

AVOGADRO'S CONSTANT

If the relative atomic mass of any element is expressed in grams and weighed, it will contain 6.02×10^{23} atoms.

For example,

Number of C atoms in 12g Carbon

„ „ S „ „ 32g Sulfur

„ „ H „ „ 1g Hydrogen, etc., are all
equal to 6.02×10^{23}

The molar mass (symbol M) is the mass of one mole of that substance. We usually use the unit g mole^{-1} . The molar mass of an element is the mass per mole. It follows from the definition of the mole that the molar mass of carbon is 12.09 mole^{-1} . Similarly, since the relative atomic mass of uranium is 238 mole^{-1} $M = 238 \text{ g mole}^{-1}$

The term molar mass applies not only to elements in the atomic state but also to all chemical species, atoms, molecules, ions, etc.

A convenient summary of these relationships can be given as :

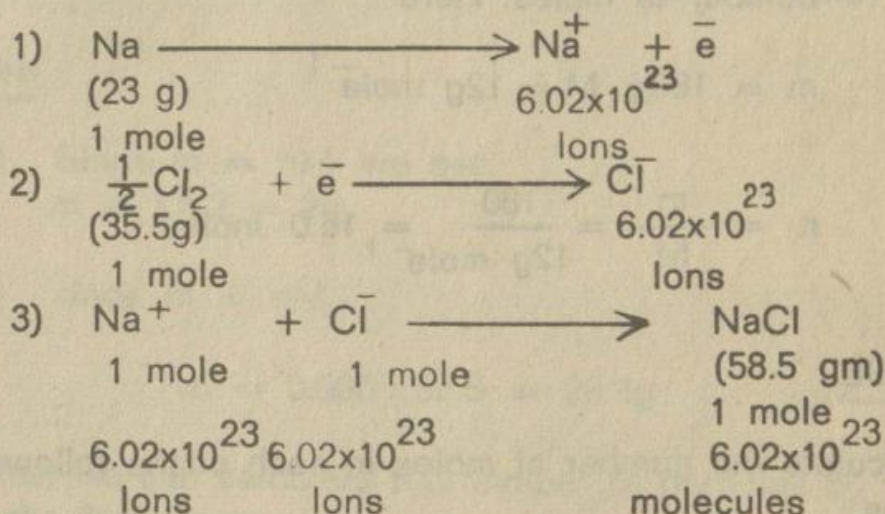
$$\text{number of moles} \times 6.02 \times 10^{23} = \text{number of particles (atoms, molecules, etc).}$$

The molecular mass is the relative mass of a molecule and is the sum of the respective atomic masses of the constituent atoms, e.g., molecular mass of water is calculated like this :

Relative mass of 2H atoms	=	2×1	=	2
„ „ 1 O atom	=	1×16	=	16
Total	=			18

This molecular mass when expressed in grams gives the mass of one mole of molecules of the substance.

Similarly for NaCl, the gram formula weight is 58.5g, i.e., 58.5g of NaCl = 1 mole of NaCl. One mole of NaCl contains 6.02×10^{23} Na^+ ions and 6.02×10^{23} Cl^- ions. Consider the following examples :



1 Mole of any gas at standard temperature and pressure has a volume of 22.4 litres and contains 6.02×10^{23} molecules of that gas. That is, 1 mole of hydrogen gas means that 2 g of hydrogen gas at S.T.P. contain 6.02×10^{23} molecules of hydrogen gas and it has a volume of 22.4 l.

OBJECTIVE EXERCISE

For calculating the amount of a substance, we have a general formula :

$$\text{Amount of substance (in moles)} = \frac{\text{mass of substance in g}}{\text{molar mass (g mole)}}$$

$$\text{or, } n = \frac{m}{M}$$

Example

A sample of carbon weighs 180g. What amount of carbon in moles is present in this sample?

Solution We can use $n = \frac{m}{M}$ which gives the desired number of moles. Here

$$m = 180\text{g}, M = 12\text{g mole}^{-1}$$

$$n = \frac{m}{M} = \frac{180}{12\text{g mole}^{-1}} = 15.0 \text{ moles}$$

EXERCISE

Calculate the number of moles in each of the following examples :

1) 50.0g of CaCO_3

$$n = \frac{m}{M} = \frac{50.0\text{g}}{100 \text{ g mole}^{-1}} = 0.50 \text{ moles}$$

2) 30 g of oxygen molecules

$$n = \frac{m}{M} = \frac{30\text{g}}{32\text{g mole}^{-1}} = 0.953 \text{ moles}$$

Exercise

Calculate the weight of substance in each of the following :

- a) 1.00 mole of hydrogen (H_2); (b) 0.500 M NaCl

Solution

- a) Since $m = nM$, we get
 $m = 1 \times 2 = 2g$

- b) since $m = nM$

$$m = 0.500 \times 58.5 = 29.3g$$

Now, we can calculate the number of particles by applying the formula ;

$$N = nL$$

where N = number of particles

n = the amount of substance (in moles) and

L = Avogadro's constant.

Exercise

Calculate the number of atoms in 18.0g of carbon.

Solution

We shall first calculate the number of moles

n by applying the formula $n = \frac{m}{M}$

$$n = \frac{18.0g}{12g \text{ mole}^{-1}} = 1.5 \text{ mole}$$

Now applying the formula : $N = nL$, we get

$$N = 1.5 \times 6.02 \times 10^{23} = 9.03 \times 10^{23} \text{ atoms}$$

EXAMPLES FROM ELECTRO-CHEMISTRY

We can use the mole concept in electro-chemistry as well. One mole of electrons will deposit one mole of Ag(silver) metal (108g) from a solution containing Ag ions.



CONCENTRATION OF SOLUTIONS

We can calculate the amount of a substance in a given volume of solution of known strength by using the formula :

$$C = \frac{n}{V}$$

where C = concentration (mole l^{-1})

V = Volume (l) and

n = number of moles.

Example 1

Calculate the concentration of a solution which is made by dissolving 0.500 mole of NaOH in 200 ml of a solution.

$$\text{Using } C = \frac{n}{V}$$

where $n = 0.500$ mole and $V = \frac{(200)}{(1000)} \text{ l}$

$$\text{we get: } C = \frac{n}{V} = \frac{0.500 \text{ mole}}{0.200 \text{ l}} = 2.50 \text{ mole l}^{-1}$$

Example 2

Assuming that 0.100 mole of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ is placed in a volumetric flask and is diluted to a volume of 2000 ml, calculate the concentration of the solution.

$$\text{Here, } V = 2000 \text{ ml} = \frac{2000}{1000} \text{ l} = 2 \text{ l}$$

$$n = 0.100 \text{ mole}$$

$$C = \frac{n}{V} = \frac{0.100 \text{ mole}}{2.00 \text{ l}} = 0.0500 \text{ mole l}^{-1}$$

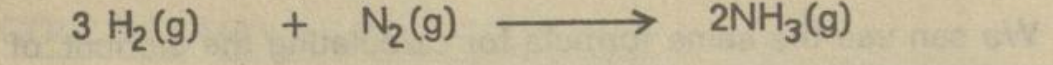
We can use the same formula for calculating the amount of a substance in moles dissolved in a solution

1) 4.00 l of 5.00 M NaOH will have :

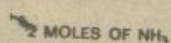
$$n = CV$$

$$n = 5.00 \text{ mole l}^{-1} \times V = 4.00 \text{ l}$$

$$n = 5.00 \text{ mole l}^{-1} \times 4.00 \text{ l} = 20 \text{ moles.}$$



2 moles

$$2 \times 6.02 \times 10^{23}$$
 molecules
$$PV = nRT$$

$$PV = nRT$$

Where R is a constant for all gases, n the amount of the gas expressed in terms of the number of molecules present. On the carbon -12 scale of relative atomic mass, one mole contains 6.02×10^{23} elementary entities. Using the equation $PV=nRT$, it is possible to calculate the volume which one mole of a gas will occupy at S.T.P. (as 22.4 litres.)

Exercise 1

A sample of ammonia weighs 1.00 g. How many moles of ammonia are contained in this sample.

$$n = \frac{m}{M}$$

$$\text{where } m = 1.0\text{g}, M = 17.0\text{g mole}^{-1}$$

$$\begin{aligned}\text{So } n &= \frac{m}{M} = \frac{1.00\text{g}}{17.0\text{g}} \text{ mole}^{-1} \\ &= 0.0588 \text{ mole}\end{aligned}$$

Exercise 2

What mass of SO_2 contains the same number of molecules as are in 1.00 g of NH_3

If the number of molecules of each gas is to be the same then the amount must be the same. For SO_2 , $M=64.1\text{g mol}^{-1}$. from the above exercise we know that $n=0.0588 \text{ mol}$.

$$\begin{aligned}\text{Substituting these values in the expression: } m &= nM \\ \text{we get } m &= nM = 0.0588 \text{ mol} \times 64.1\text{g mol}^{-1} \\ &= 3.77 \text{ g.}\end{aligned}$$

NOTE

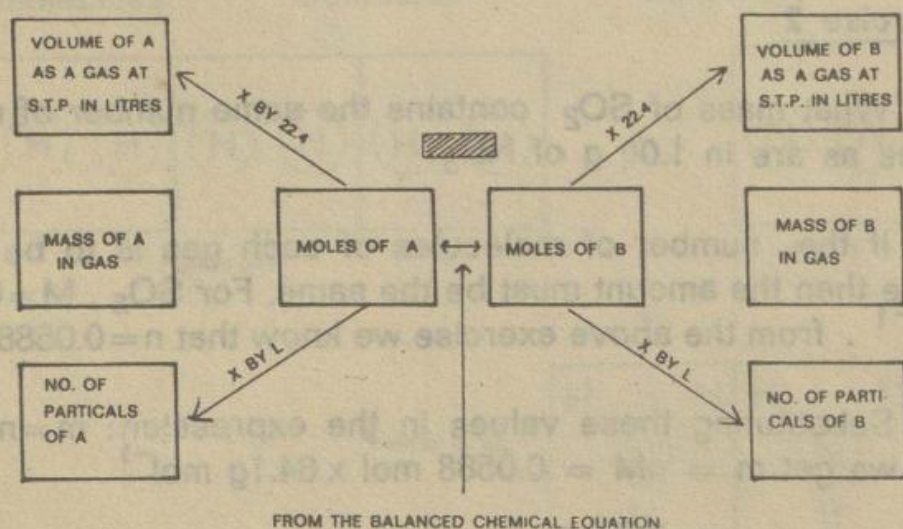
When giving examples to pupils it is advisable to use substances whose formula masses are whole numbers.

Here is the list of few substances whose formula masses are whole numbers.

Substance	Relative formula mass
HF	20
Li_2O , C_2H_6	30
Ca, MgO, KH, NaOH	40
SiO_2 , $\text{C}_3\text{H}_7\text{OH}$, CH_3COOH	60
CuO , NH_4NO_3 , SO_3	80
CaCO_3 , KHCO_3 , Mg_3N_2	100
Al_2S_3 , NaI	150
CaBr_2 , Cr_2S_3 , Hg	200

With the help of this list, the problem of finding percentage composition and empirical formula becomes easier.

The following diagram indicates many relationships between the mole and other areas in chemistry.



From a balanced equation the relative number of moles of reactants and products can be determined. The amount of A and B can be expressed in terms of the number of particles involved, in grams or in litres if gases. This is

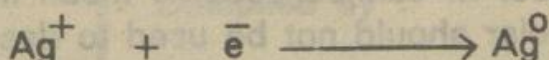
obtained by multiplying the number of moles shown in the equation by the appropriate conversion factor, 22.4 Mr (Mr is now the standard abbreviation for relative molecular mass).

The term molar mass applies not only to elements in the atomic state but also to all chemical species, atoms, molecules, ions etc.

The unit for amount is mole, for mass is gram, for molar mass is gram per mole.

Faraday's Law

For electrochemical processes, the problems involving quantitative measurements can be solved using the concept of mole. The equation is written for each electrochemical process and the Faraday is taken as the quantity of electricity associated with an Avogadro number of electrons. For example, the reaction



involves 1 mole of Ag^+ ions combining with 1 mole of $\bar{e}$ (electrons) to produce Ag metal.

In addition, the mole concept can be applied to problems based on the equilibrium law, osmosis, solubility products, etc. The concept of gram molecular weights or equivalent weights are confusing for the students. The difficulty thus encountered can be reduced when the concept of mole is used instead.

Molar volume of Solutions

When two solutions react together there is usually a simple ratio between the number of moles of reactants. For

this reason, it is convenient to represent the concentration of the solute in a solution in terms of the number of moles of it in a particular volume of solution. Thus chemists measure the concentration as moles per litre. For example, a solution of HCl (aq) containing 1.00 mole per litre will have 36.5 g of HCl in one litre of solution. A solution with 2.00 moles of HCl per litre contains 73g ($2 \times 36.5\text{g}$) in 1 litre of solution.

These concentrations are expressed as the number of moles per litre of the solution and not 1 litre of the solvent.

At one time, a solution containing 1.0 mole of solute per l was described as "one molar" but the terms molar and molarity are now obsolete and it is wrong to use them in the context of solution concentrations. The units associated with the word "molar" are "per mole" so it is quite incorrect to refer to a solution containing 0.1 mole of solute per l as 0.1 molar. Nevertheless, 0.1 M for 0.1 moles per litre is still used. Thus the symbol M can be used to mean moles per litre but the word molar should not be used to describe the concentrations of solutions.

**EXPERIMENTS
FOR
CLASS IX & X CHEMISTRY
COURSE**

Compiled by :

Dr. M. Jaffar,
Quaid-i-Azam University,
Islamabad.

EXPERIMENT 1 To estimate the size of atoms/ions

RELEVANCE TO TEXT 2 · 6, 2 · 9, 5 · 1, 5 · 6

PROCEDURE

1. Dissolve 1 g of KMnO_4 per litre of its solution in H_2O .
2. Using a measuring cylinder take 10 ml of the solution and place it in a 100-ml beaker.
3. Calculate the mass of KMnO_4 in the beaker.
4. Pour water into the beaker upto the mark and Stir.
5. After this dilution, the mass of KMnO_4 remains the same in the beaker. Now, take out 10 ml portion of this diluted solution and dilute it 10 times further in another 100-ml beaker.
6. Calculate the amount of KMnO_4 in the resulting solution.
7. Continue diluting the KMnO_4 solution until its colour is just only visible.
8. Calculate the mass of KMnO_4 in solution before the colour disappears.

OBSERVATIONS

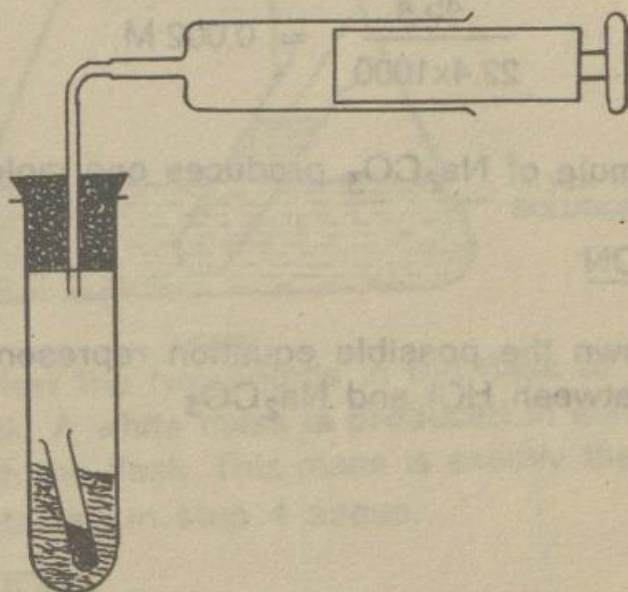
1. How does the final amount of KMnO_4 in solution furnish information on the size of MnO_4^- ion responsible for producing colour in solution?
2. How sizes of ions are related to sizes of atoms?

EXPERIMENT 2 To establish equation for the reaction between sodium carbonate and hydrochloric acid.

RELEVANCE TO TEXT 2 · 20, 15 · 7, 7 · 10.

PROCEDURE

1. Weigh accurately 0.2 g Na_2CO_3 (anhydrous) and place it in a small test tube.
2. Carefully lower this test tube in a big test tube containing about 20 ml 5M HCl.
3. Connect the large test tube to a gas syringe, of capacity 100 ml, by a rightangled glass tube. See Figure.
4. Note the reading on the gas syringe.
5. Tip the large test tube so that the acid mixes thoroughly with the carbonate and CO_2 is evolved.
6. When evolution of the gas stops, read the syringe.
7. Note the volume of CO_2 collected.



FIGURE

SPECIMEN CALCULATION

Mass of anhydrous sodium carbonate = 0.20 g

Volume of CO_2 evolved = 47 ml

Room temperature = 10°C

$$\begin{aligned} \text{a) Number of moles of sodium carbonate taken} &= \frac{0.20}{106} \\ &= 0.002 \text{ M} \end{aligned}$$

b) Number of moles of CO_2 evolved :

Reduce volume of CO_2 to S.T.P.

The volume of CO_2 at 0°C will be

$$\frac{47 \times 273}{283} = 45.4 \text{ ml}$$

Now 22.4 litre of any gas at S.T.P. contains one mole of gas.
Hence, 45.4 ml of CO_2 at S.T.P. contains.

$$\frac{45.4}{22.4 \times 1000} = 0.002 \text{ M}$$

Hence, one mole of Na_2CO_3 produces one mole of CO_2

OBSERVATION

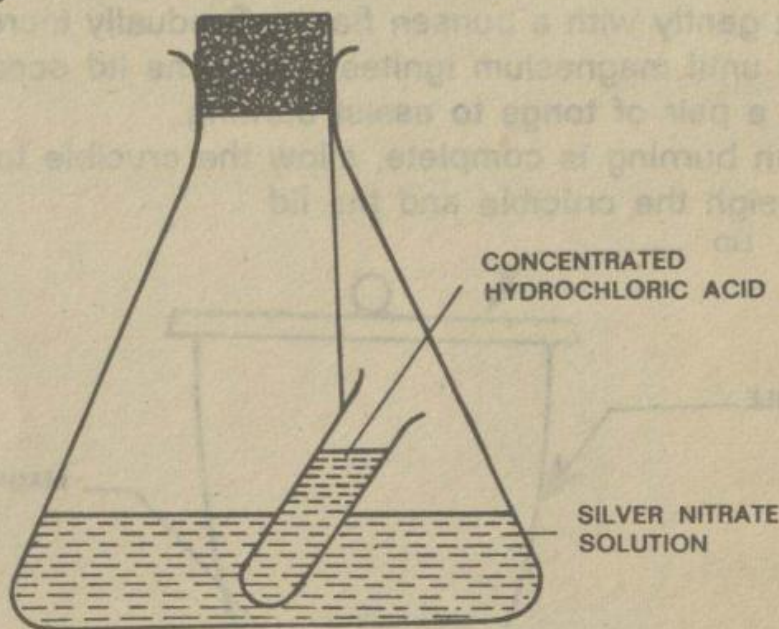
1. Write down the possible equation representing the reaction between HCl and Na_2CO_3

EXPERIMENT 3 To prove the Law of Conservation of Mass

RELEVANCE TO TEXT 2 · 14

PROCEDURE

1. Put some silver nitrate solution into a conical flask (250-ml capacity or smaller).
2. Lower into the flask by means of a thread a small test tube containing some hydrochloric acid.
3. Insert the stopper as shown in Figure.
4. Weigh the flask on a pan of a balance.



FIGURE

5. Now allow the two liquids in the flask to mix by tilting the flask. A white mass is produced in solution.
6. Reweigh the flask. This mass is exactly the same as the one obtained in step 4 above.

OBSERVATIONS

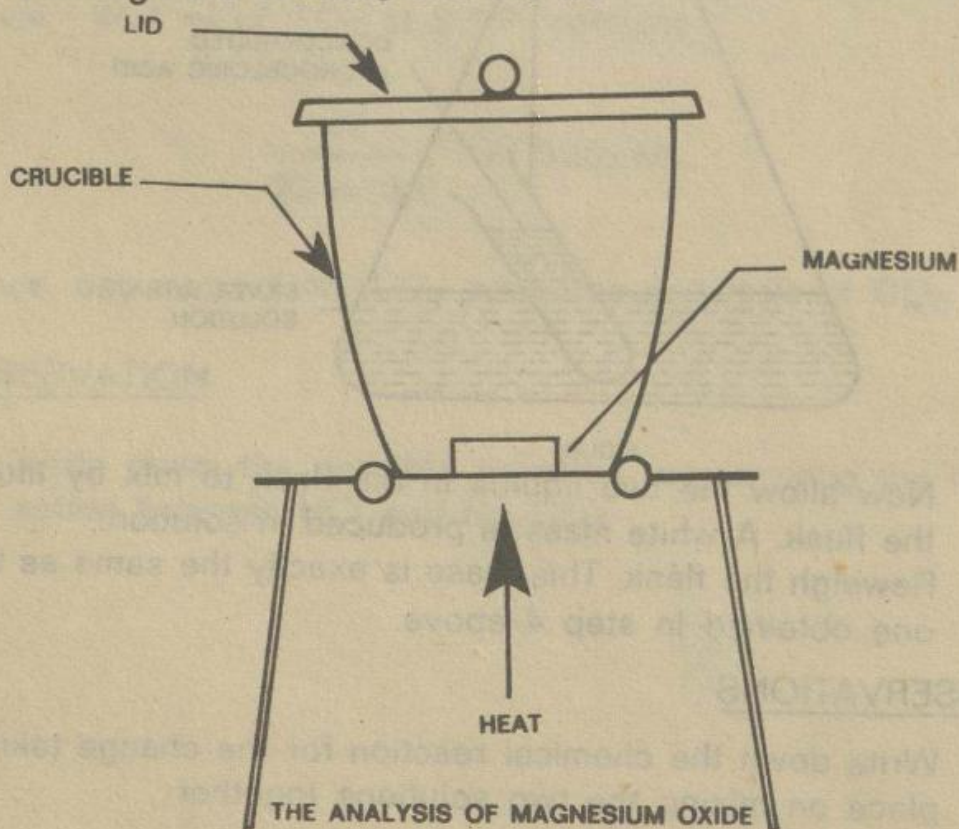
1. Write down the chemical reaction for the change taking place on mixing the two solutions together.
2. Infer results from the experiment.

EXPERIMENT 4 To determine the formula of magnesium oxide.

RELEVANCE TO TEXT 2 · 19

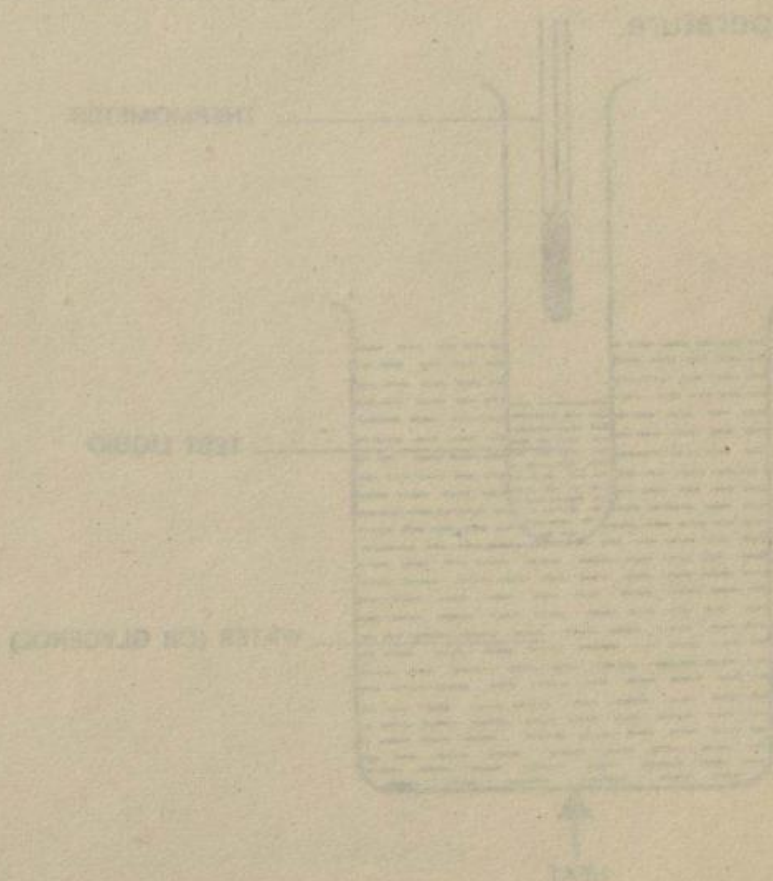
PROCEDURE

1. Weigh a crucible along with its lid.
2. Weigh into the crucible 0.24 g of clean magnesium ribbon.
3. Place the crucible (with its lid on it) on a pipe-clay triangle placed on a tripod. See Figure.
4. Heat gently with a bunsen flame. Gradually increase the heat until magnesium ignites, lifting the lid occasionally with a pair of tongs to assist burning.
5. When burning is complete, allow the crucible to cool.
6. Reweigh the crucible and the lid.



OBSERVATIONS

1. What fraction of a mole of magnesium atoms is 0.24 g ?
2. What mass of oxygen combines with 0.24 g of magnesium ?
3. Calculate fraction of a mole of oxygen atoms combining with 0.24 g of Mg and hence find the formula of magnesium oxide.

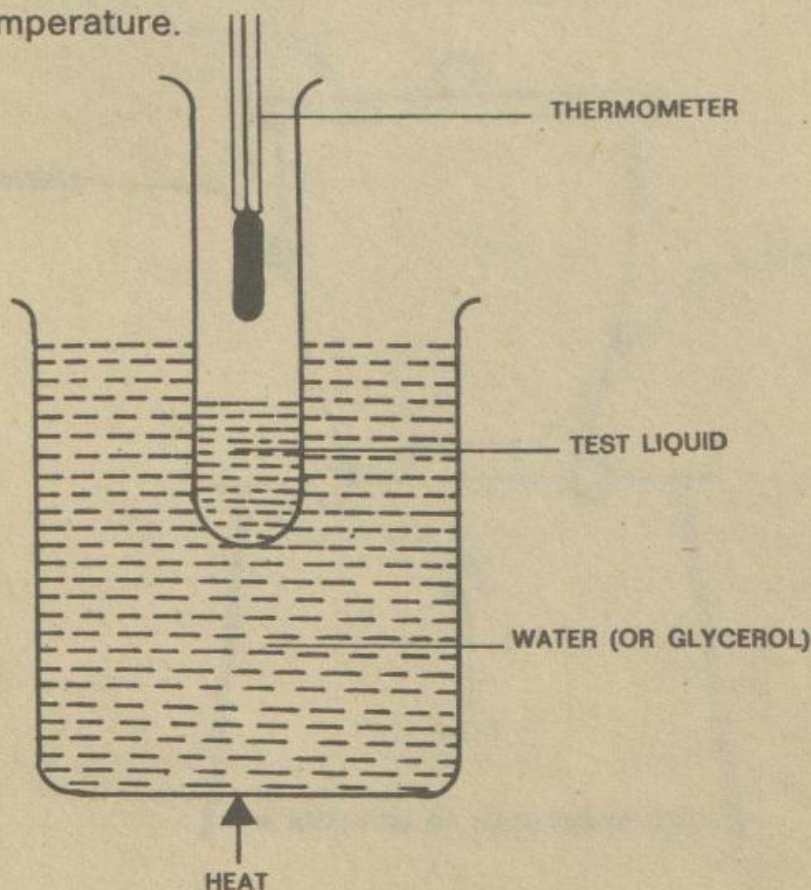


EXPERIMENT 5 To determine the boiling point of a liquid.

RELEVANCE TO TEXT 4 · 15

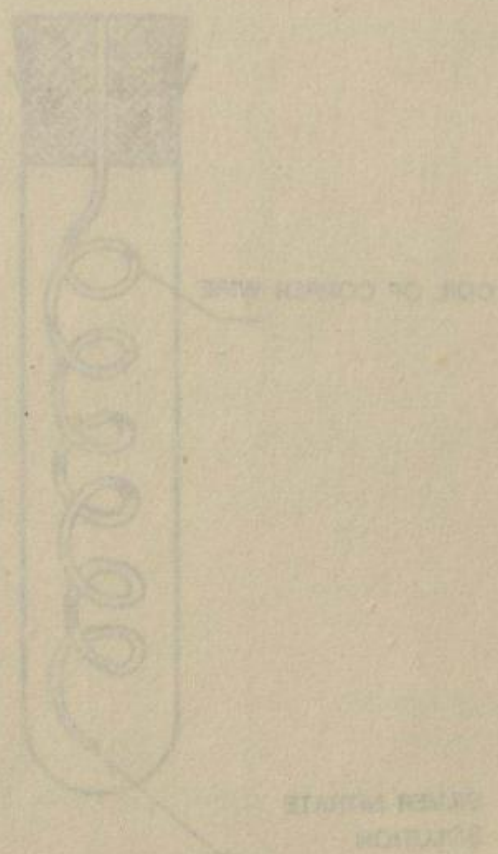
PROCEDURE

1. Take 5—10 ml of ethanol in a small test tube.
2. In a 250-ml beaker, pour about 200 ml glycerol or water, and place this beaker on a tripod.
3. Clamp the tube from a stand such that its lower part is fully immersed in water (or glycerol).
4. Also, hang a thermometer (0—100°C) right into the tube with the precaution that its lower end should remain above the surface of the ethanol.
5. Start heating the beaker, and go on noting the temperature increase.
6. When the ethanol starts boiling, note the final steady temperature.



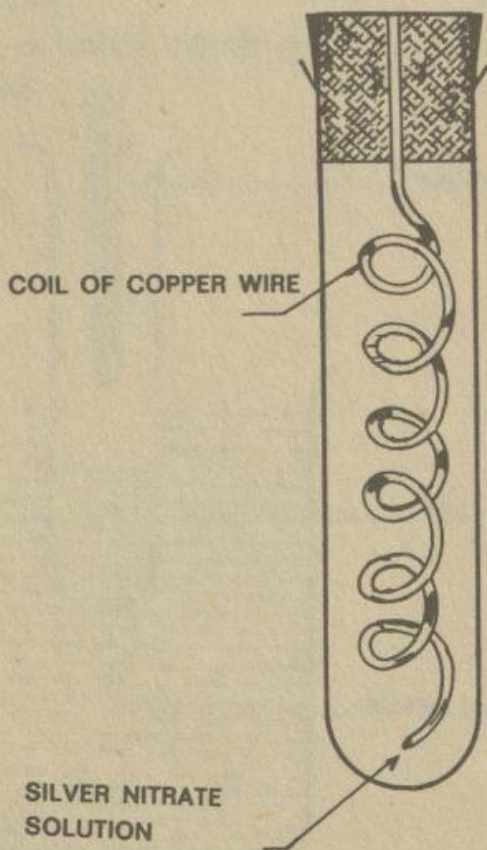
OBSERVATIONS

1. Why glycerol is also recommended in this experiment ?
2. What is the disadvantage of immersing the bulb of the thermometer into the liquid ?
3. To which liquids is the above method applicable ?



EXPERIMENT 6 To show that metals are crystallineRELEVANCE TO TEXT 4 · 16, 4 · 17PROCEDURE

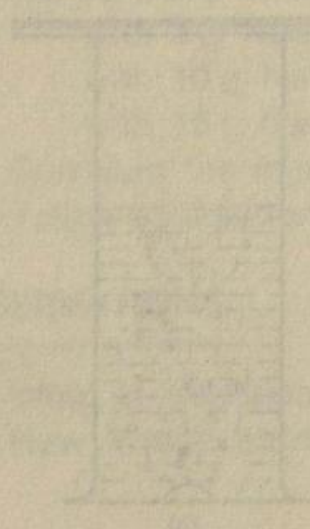
1. Take about 15 ml silver nitrate solution in a test-tube.
2. Make a coil of copper wire as shown in Figure and pass this coil through a cork fitted in the mouth of the test-tube.
3. Allow the coil to stand in solution for 15 minutes.
4. Observe what has happened to the surface of the coil.
5. Examine the crystals of silver metal accumulated on copper coil under a lens.



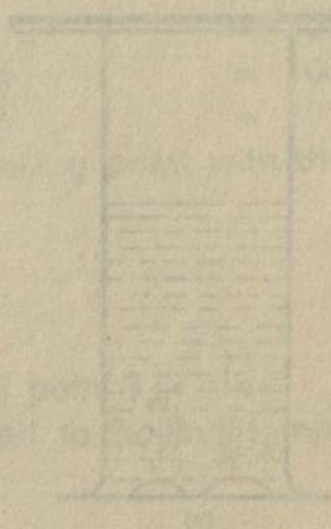
PREPARING CRYSTALS OF METALS

OBSERVATIONS

1. What is the reaction responsible for the production of silver metal?
2. Repeat experiment with a zinc foil immersed in lead nitrate solution and infer what happens to zinc foil.



(a)



(b)

OBSERVATIONS

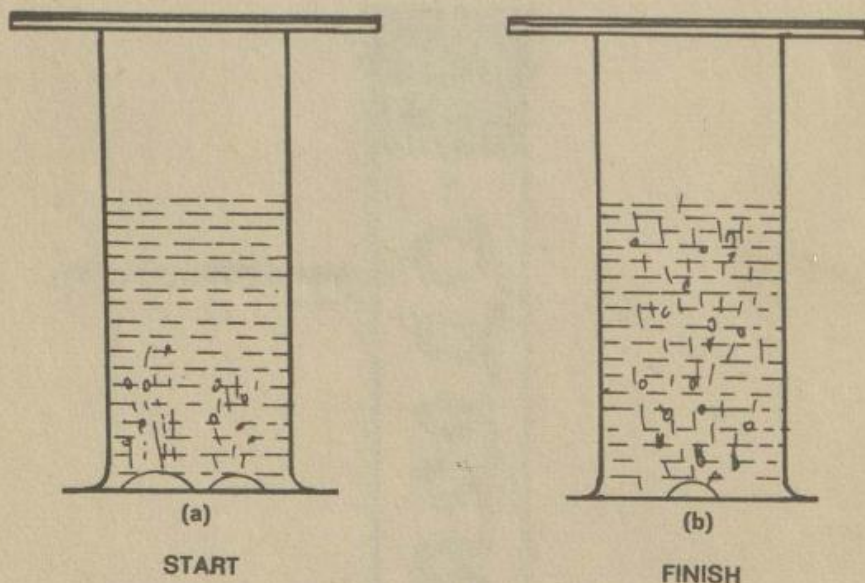
1. How diffusion is explained on the basis of this experiment?
2. How this experiment illustrates the fact that matter is composed of tiny particles?

EXPERIMENT 7 To demonstrate the diffusion of coloured crystals in solution.

RELEVANCE TO TEXT 5 · 3

PROCEDURE

1. Half fill a 250-ml jar or conical flask with water.
2. Drop a crystal of a deeply coloured salt (such as potassium permanganate or ammonium dichromate) into water as shown in Figure.
3. During a period of 10-15 minutes observe closely what happens within the flask or jar.



OBSERVATIONS

1. How diffusion is explained on the basis of this experiment ?
2. How this experiment illustrates the fact that matter is composed of tiny particles ?

EXPERIMENT 8 To study the effect of dissolved solids on the boiling point of a solution.

RELEVANCE TO TEXT 5 · 4, 5 · 6

PROCEDURE

1. Arrange apparatus as for experiment on the determination of boiling point of a liquid.
2. Prepare different solutions of NaCl in water by dissolving 5 g, 10 g and 15 g of NaCl in 150 ml water.
3. Determine the boiling point in each case using a glycerol bath.
4. Let the data obtained be of this type

Temperature of boiling water	= 99.7°C
with 5 g NaCl solution	= 100.2°C
with 10 g NaCl solution	= 100.7°C
with 15 g NaCl solution	= 101.2°C
5. Correlate the increase in boiling point with the concentration of solution.

OBSERVATIONS

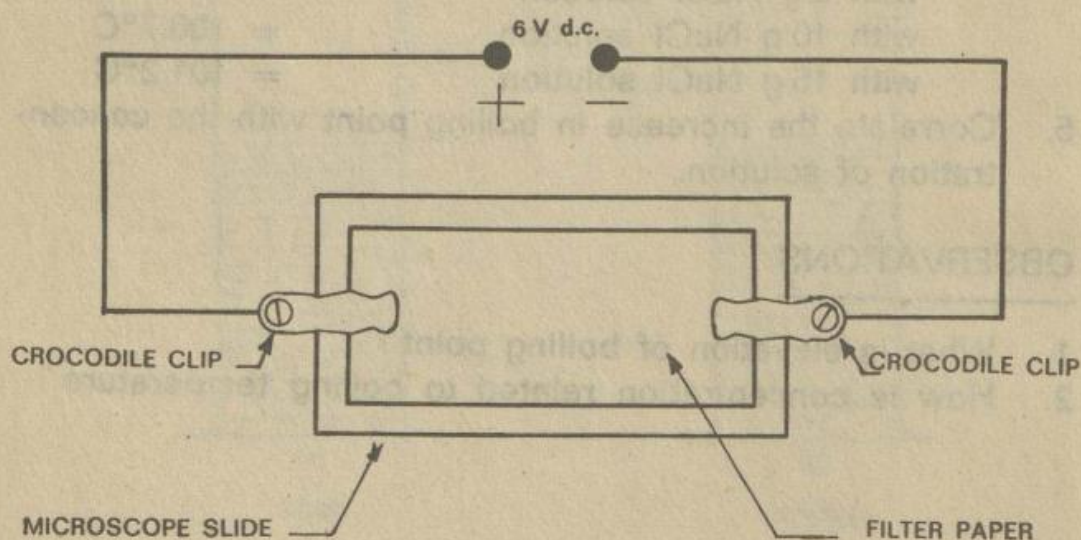
1. What is elevation of boiling point?
2. How is concentration related to boiling temperature?

EXPERIMENT 9 To observe the movement of ions.

RELEVANCE TO TEXT 6.3,

PROCEDURE

1. Fold a piece of filter paper several times to form a thick pad, approximately 1 cm in width.
2. Saturate the pad with tap-water.
3. Fasten the pad to a microscope slide by means of two crocodile clips.
4. Connect the circuit as shown in Figure.
5. Place a small crystal of potassium permanganate in the centre of the filter paper pad, and leave with the current on for about 15 minutes.
6. Observe the movement of a coloured zone towards one side from the centre of the pad.



THE MOVEMENT OF IONS

OBSERVATIONS

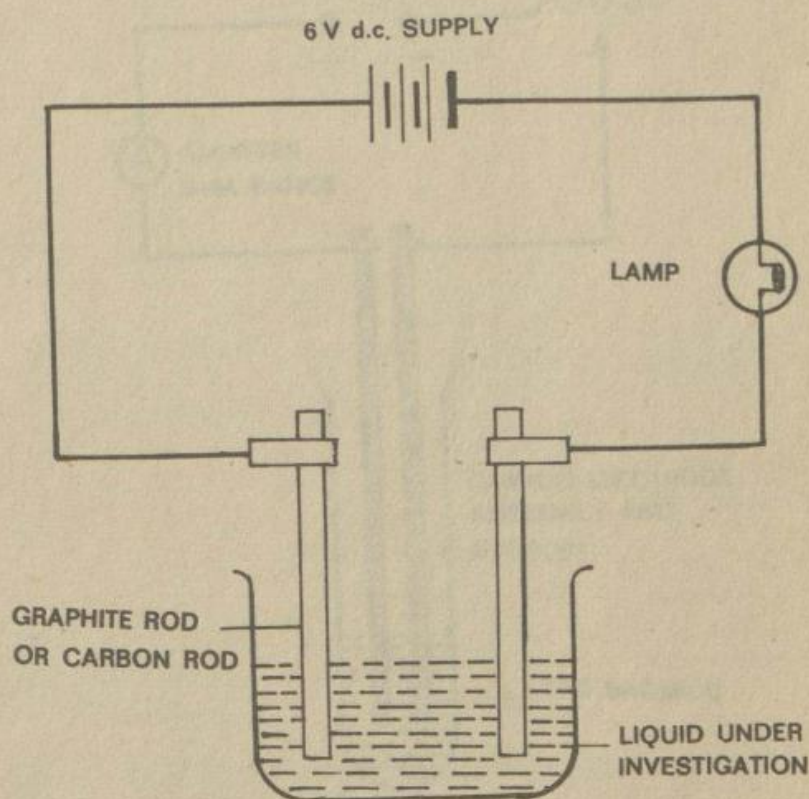
1. Why is the coloured zone movement to one side ?
2. What this colour is due to ?

EXPERIMENT 10 To study the conduction of electricity by a variety of substances in the solid and liquid states.

RELEVANCE TO TEXT 6 · 3,

PROCEDURE

1. Arrange the circuit shown in Figure.
2. First place a dilute solution of NaCl in the beaker and see if the lamp glows on immersing the electrodes into the beaker.
3. Then repeat step 2 above, for pure water, alcohol, turpentine and benzene turn by turn.

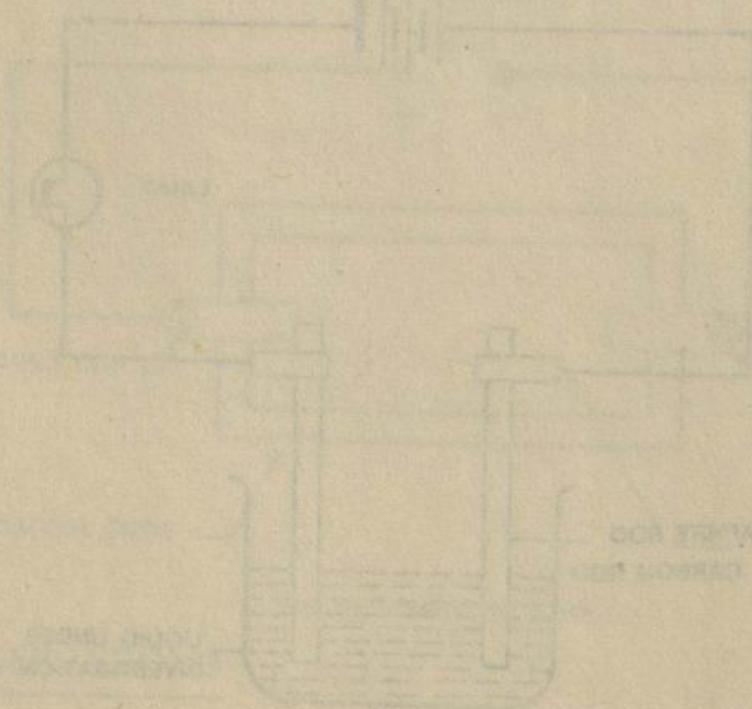


FIGURE

4. Small pieces of solids can be tested for electrical conduction also. Try lead, copper, rubber, plastic, paper salt, sugar, magnesium, etc., by pressing the electrodes against them.
5. Classify these liquids and solids as conductors and non-conductors.

OBSERVATIONS

Tabulate your findings.

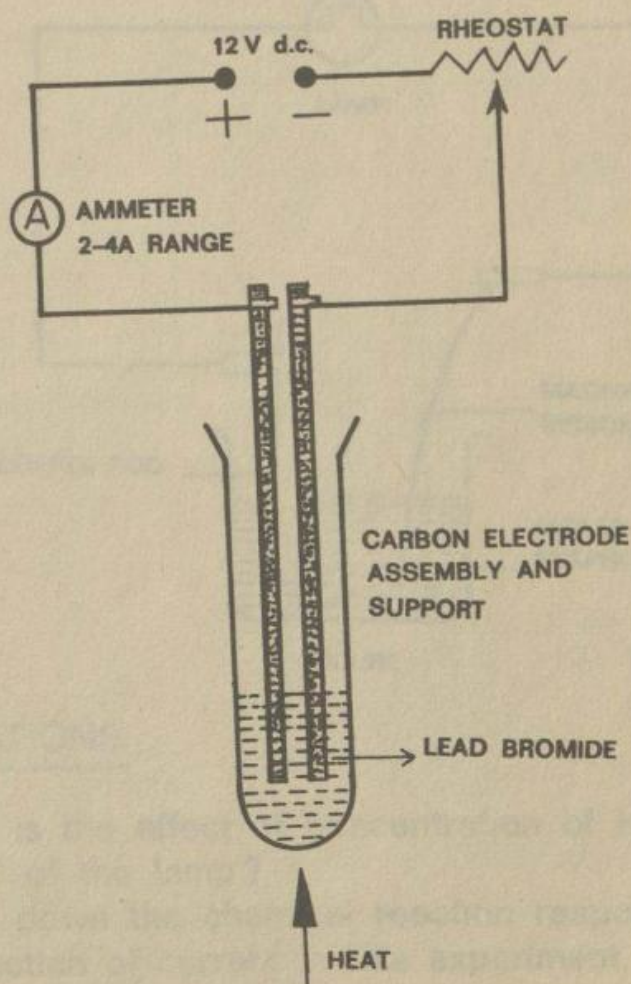


EXPERIMENT 11 To determine the quantity of electricity needed to deposit one mole of lead atoms.

RELEVANCE TO TEXT 6 . 5, 6 . 6.

PROCEDURE

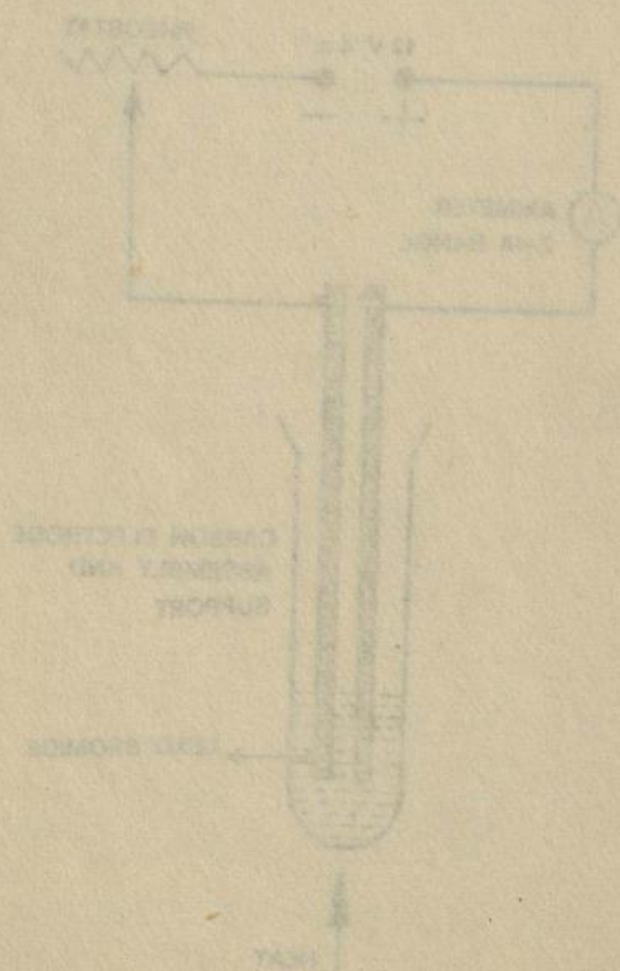
1. Set up the apparatus as shown in Figure. A 6-V d.c. supply is satisfactory if 12 V d.c. is not available.
2. Heat the electrolyte until completely molten.
3. Adjust current at a constant value within 2 to 4 A, start a stop-clock simultaneously.
4. Run the steady current through the circuit for 15 minutes.



5. Switch off the current and stop the clock.
6. Pour off the molten lead bromide and the lead bead formed onto a clean dry dish. Separate the lead bead and weigh it.

OBSERVATIONS

1. Calculate the number of coulombs passed that gave the mass of lead weighed.
2. Calculate the number of coulombs required to deposit one mole of lead atoms. (207 g).

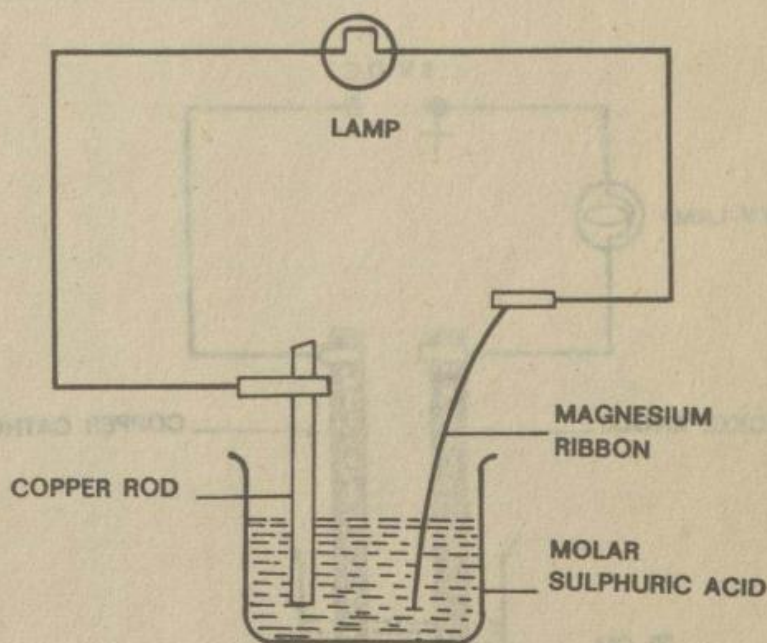


EXPERIMENT 12 To produce electric current from a chemical reaction.

RELEVANCE TO TEXT 6 - 7

PROCEDURE

1. Set up the circuit shown in Figure, using a 1.5V lamp.
2. Immerse the electrodes in 1M H_2SO_4 contained in a large beaker.
3. Observe the bulb which will glow as a result of electric current produced because of a chemical reaction between Cu/Mg and H_2SO_4 .



OBSERVATIONS

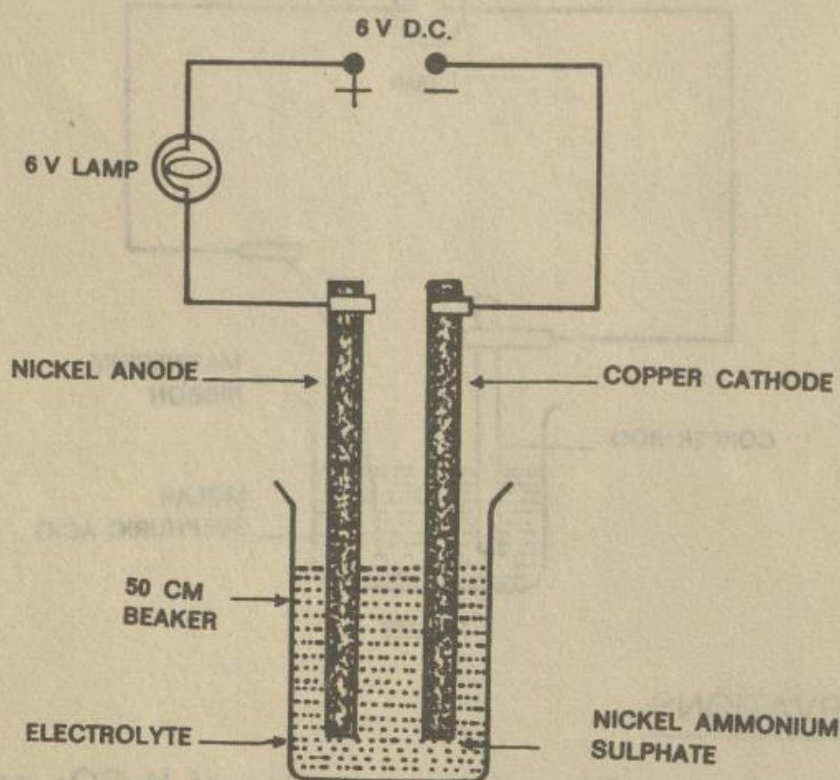
1. What is the effect of concentration of H_2SO_4 on the 'glow' of the lamp?
2. Write down the chemical reaction responsible for the production of current in this experiment.

EXPERIMENT 13 To demonstrate nickel-plating.

RELEVANCE TO TEXT 6 - 11.

PROCEDURE

1. Set up the apparatus as shown in Figure.
2. Clean and dry the copper cathode before immersing it in nickel ammonium sulphate solution.
3. Switch on the current for 10 minutes.
4. Switch off the current, and remove the cathode.
5. Rinse the cathode with water, and allow to dry. It has been nickel plated.

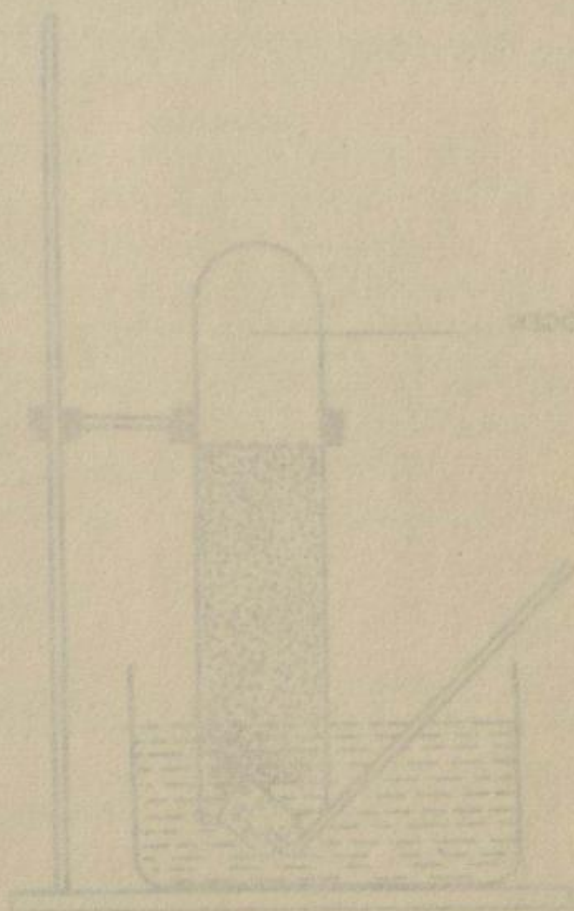


ELECTROPLATING WITH NICKEL

FIGURE

OBSERVATIONS

1. What other metals can be electroplated by this method ?
2. What is the use of including a lamp in series in the circuit ?
3. Why is it necessary to have a neat, dry surface before electroplating ?



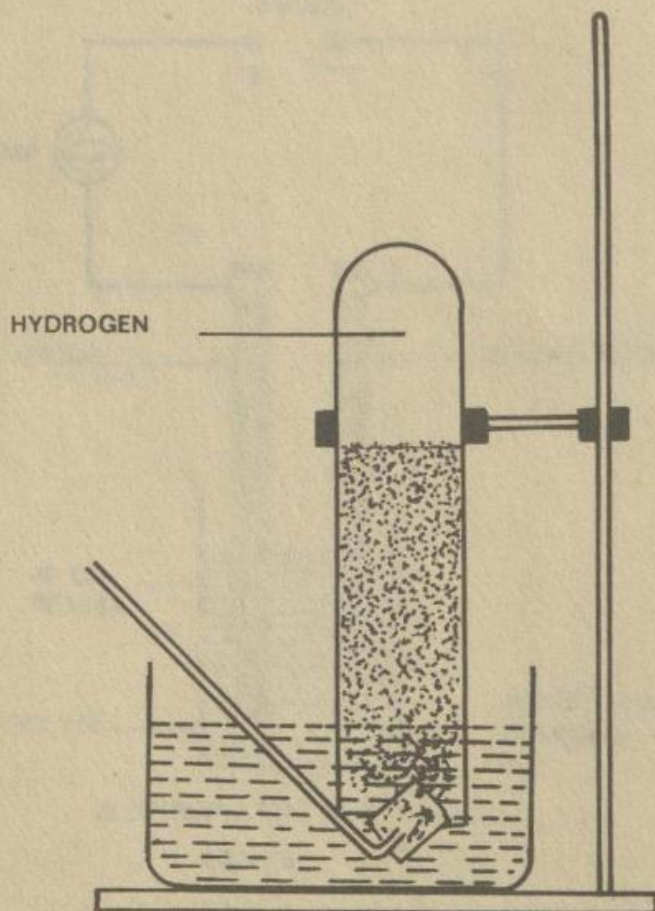
EXPERIMENT 14 To study reaction between hydrochloric acid and magnesium.

RELEVANCE TO TEXT 7 · 10, 15 · 7, 2 · 20.

NOTE This experiment is to be performed by the teacher only.

PROCEDURE

1. Place in a large trough enough 1–2 N hydrochloric acid and arrange the apparatus as shown in figure.



FIGURE

2. Place a small piece of magnesium metal in the trough and immediately lower the large boiling tube (filled with water) over it.
3. Clamp the tube in position, and collect the gas given off.

OBSERVATIONS

1. What is the reaction between HCl and Mg metal?
2. How the gas will be checked to be H_2 ?

OBSERVATIONS

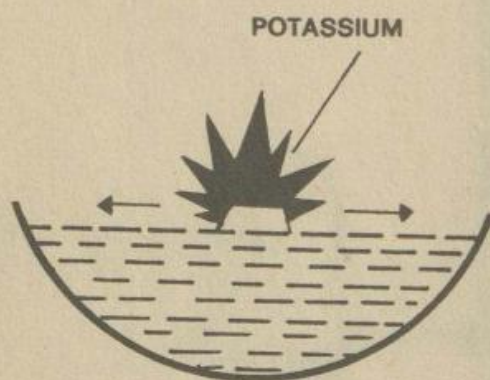
OBSERVATIONS

1. What gas is given off when potassium reacts with water? How to check it?
2. What is the chemical reaction between potassium and water?

NOTE - Goggles must be worn for this experiment.

EXPERIMENT 15 To study the action of potassium on water**RELEVANCE TO TEXT** 7. 6**PROCEDURE**

1. Cut a small piece of potassium metal about 2 to 3 mm cube in size.
2. Place this piece of potassium on water contained in a large dish. See Figure.
3. Notice what happens to potassium, and if a gas is given off.
4. When the reaction is complete, test the resulting solution with a few drops of phenolphthalein indicator.



FIGURE

OBSERVATIONS

1. What gas is given off when potassium reacts with water? How to check it?
2. What is the chemical reaction between potassium and water?

NOTE Goggles must be worn for this experiment.

EXPERIMENT 16 To prepare exactly 0.1M sodium carbonate solution.

RELEVANCE TO TEXT 7 - 8

PROCEDURE

1. Clean and dry a 250-ml beaker.
2. Weigh the beaker, and in it accurately weigh 2.65 g of anhydrous Na_2CO_3 . [$\text{Na}_2\text{CO}_3 = 106$, hence 0.1M solution will require 10.6 g of Na_2CO_3 per litre or 2.65 g per 250 ml of solution].
3. Pour about 100 ml of distilled water in the beaker. Put in a glass rod, and start stirring, until Na_2CO_3 has dissolved.
4. Take a graduated (250-ml) flask. Put a funnel in the neck of the flask.
5. Pour the solution of Na_2CO_3 from the beaker into the flask, guiding it with the glass rod and wiping it off with distilled water from a wash bottle.
6. Wash the beaker sides with distilled water from the wash bottle, and allow all these washings go to the flask.
7. Remove the funnel, and add distilled water from the wash bottle directly to the solution in order to bring the solution level upto the 250-ml mark.
8. Stopper the flask, and label it as 0.1M Na_2CO_3 .

OBSERVATIONS

Study the properties of the prepared solution as given in the text.

EXPERIMENT 17 To prepare and standardize 1 N solution of hydrochloric acid.

RELEVANCE TO TEXT 7 · 8

PROCEDURE

1. Measure out approximately 100 ml of concentrated hydrochloric acid in a measuring cylinder.
2. Transfer the acid to a 1-litre capacity graduated flask using a clean glass funnel.
3. Make up to the mark by adding distilled water.
4. Standardize this solution against 1 N Na_2CO_3 solution prepared previously.

Calculations

Let the volume of HCl solution used for titration be x ml.
Let the volume of 1 N Na_2CO_3 used against x ml of HCl be 25 ml.

Hence,

$$x \text{ ml approximately N HCl} = 25 \text{ ml 1 N Na}_2\text{CO}_3$$

$$\therefore 1 \text{ ml approximately N HCl} = \frac{25}{x} \text{ ml 1 N Na}_2\text{CO}_3$$

$\therefore$ The acid is $\frac{25}{x}$ times stronger than the carbonate.

$\therefore$ The acid is $\frac{25}{x}$ N since the carbonate is 1 N

Now, label the acid prepared as $\frac{25}{x}$ N HCl.

EXPERIMENT 18 To demonstrate a reversible reaction.RELEVANCE TO TEXT 7. 11PROCEDURE

1. Put about 20 ml bromine water into a 100-ml beaker.
2. Place the beaker on a sheet of white paper.
3. Using a pipette, add 2M sodium hydroxide solution slowly to the bromine water with occasional stirring.
4. Note any colour change.
5. After the colour change add drops of 2M sulphuric acid from a second pipette.
6. Note the restoration of original colour.
7. Repeat the addition of drops of 2M NaOH again. Note the reversal of colour.

OBSERVATIONS

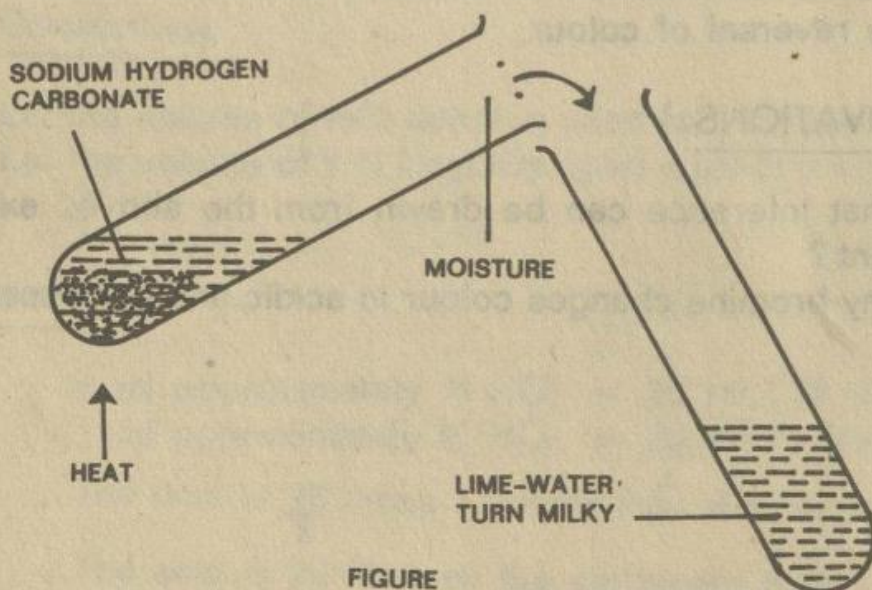
1. What inference can be drawn from the above experiment?
2. Why bromine changes colour in acidic and basic media?

EXPERIMENT 19 To study the action of heat on sodium hydrogen carbonate.

RELEVANCE TO TEXT 7 · 13, 7 · 14.

PROCEDURE

1. Place a small amount of NaHCO_3 in a dry test tube.
2. Hold the tube close enough to a large test tube so that its lip projects into the large tube which contains lime-water.
3. Heat the tube gently.
4. Observe what happens to lime-water.



FIGURE

OBSERVATIONS

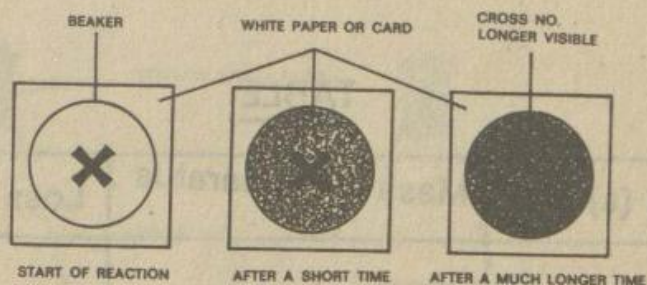
1. Write reaction that takes place on heating the hydrogen-carbonate.
2. Which gas can turn lime water milky ?
3. What is the white residue left in the tube after heating is over ?

EXPERIMENT 20 To investigate the effect of concentration on rate of reaction.

RELEVANCE TO TEXT 9.5

PROCEDURE

1. Place 50 ml of sodium thiosulphate solution (containing 40 g per litre) in a 100-ml beaker.
2. Add 5 ml 2M hydrochloric acid and at the same time start a stopclock.
3. Swirl the beaker carefully and place it on a white paper with a cross drawn on it. See Figure.
4. Observe the cross from the top of the beaker until it is no longer visible. Stop the clock at the moment.
5. Vary the concentration of thiosulphate solution by taking 40, 30, 20, 10 ml portions of original solution and mixing them separately with 10, 20, 30, 40 ml portions of distilled water in order to make a 50 ml sample in each case.
6. Repeat for each set step number 2, and record the time as per step number 4.



Figure

OBSERVATIONS

1. Draw a table showing the way in which the time taken for the cross to disappear varies with the concentration of the thiosulphate solution.
2. Infer the concentration dependence of reaction rates.

EXPERIMENT 21 → To study the effect of particle size on the rate of a reaction.

RELEVANCE TO TEXT 9 · 5, 9 · 7

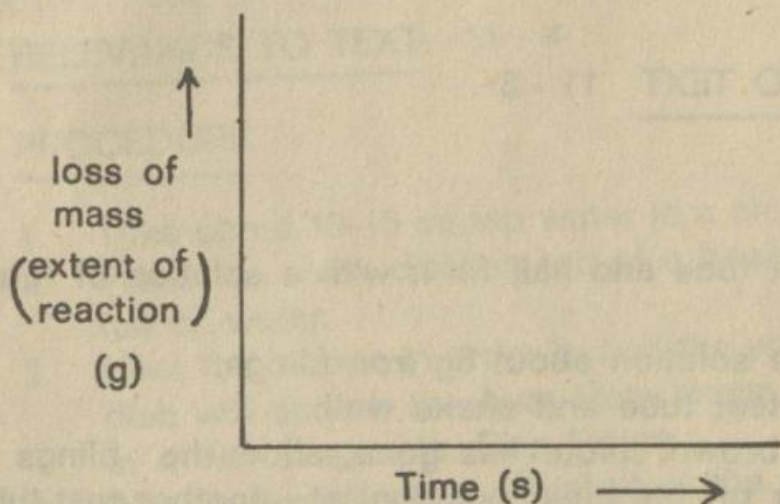
PROCEDURE

1. Grade some marble chips as large, medium and small lumps and weigh out a 20 g portion of each grade on a watch glass.
2. Weigh the watch glass with the large lumps of marble together with a 100-ml conical flask containing 40 ml 2M HCl and plugged with cotton wool, as shown in Figure (a).
3. Add the marble chips to the acid, and replace the cotton wool plug. Start the stop-clock immediately. [Remember to return all the previously weighed items to the balance pan].
4. Record the mass after thirty-second intervals, until there is no further loss of mass.
5. Tabulate results as shown in Table.

TABLE

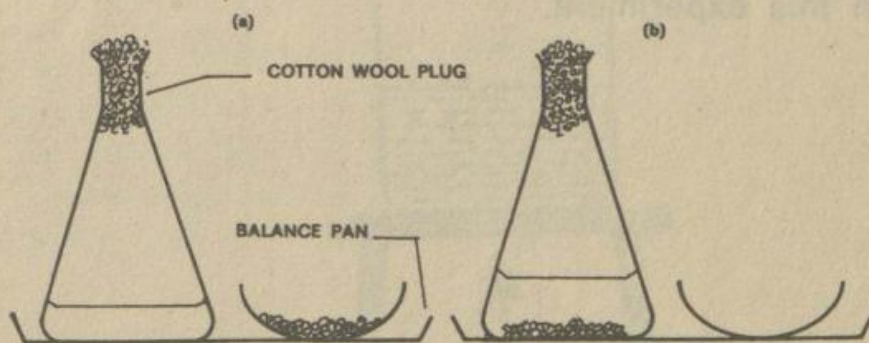
Time (s)	Mass of apparatus (g)	Loss in mass (g)
0		0
30		
60		
90		
120		

6. Repeat the experiment with other grades of chips.
7. Plot the results of the three experiments as shown below:



OBSERVATIONS

1. How is particle size related to the extent of reaction ?
2. What is the expected shape of graph between loss of mass and time in each case ?



FIGURE

THE EFFECT OF PARTICLE SIZE ON REACTION RATE

EXPERIMENT 22 To study the reduction of ferric state to ferrous state

RELEVANCE TO TEXT 11 . 3

PROCEDURE

1. Take a test tube and half fill it with a solution of ferric chloride.
2. Add to this solution about 5g iron filings.
3. Warm the test tube and shake well.
4. When the brown colour has gone, allow the filings to settle. Pour off the clear solution into another test-tube and note its colour.
5. To the clear solution add a few drops of potassium thiocyanate. What happens to the solution?

OBSERVATIONS

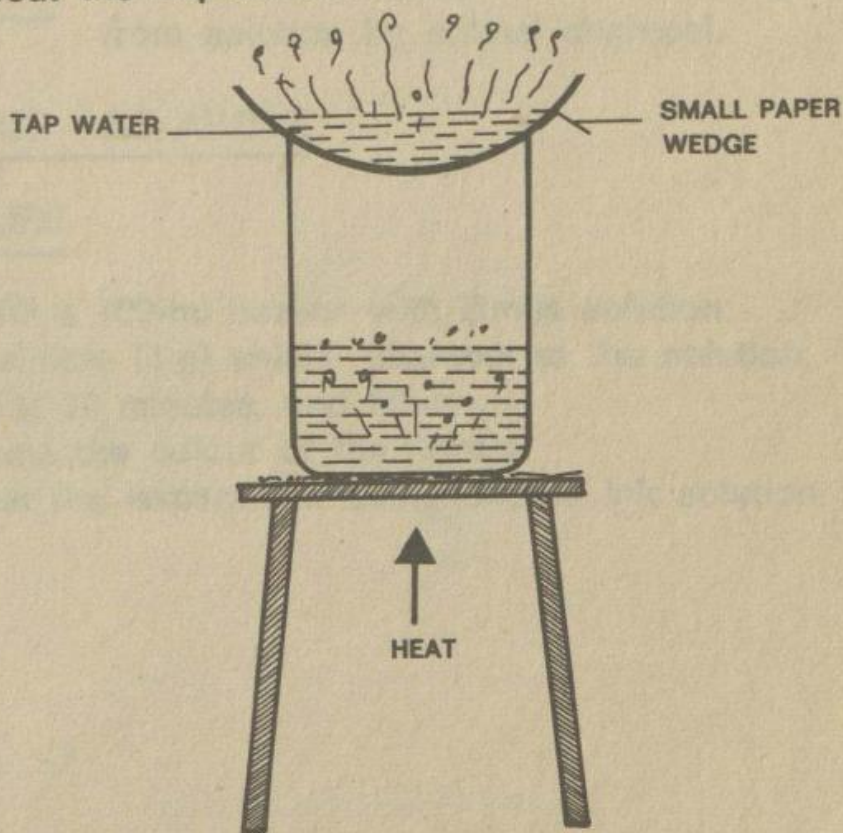
1. Why a red colour is not produced in step 5 above?
2. Write the overall chemical change that has taken place in this experiment.

EXPERIMENT 23 To show tap water contains dissolved solids.

RELEVANCE TO TEXT 11 · 4

PROCEDURE

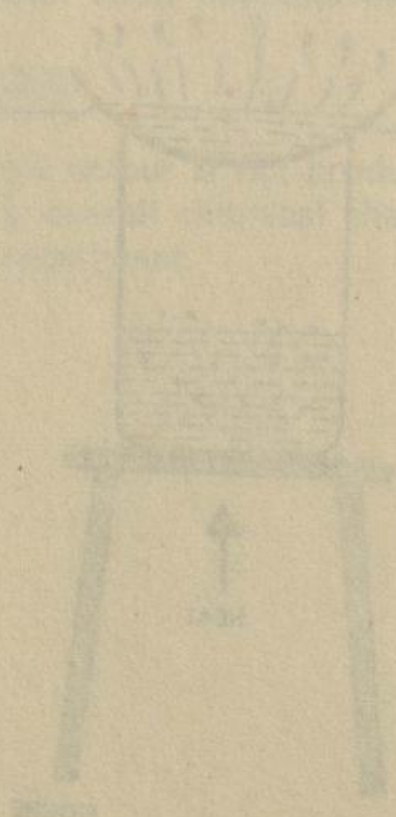
1. Take about 10-15 ml tap water in a clean, dry china dish.
2. Place the china dish on top of a beaker about one-third full of water.
3. Heat the beaker in order to boil the water in it. The china dish will absorb the heat from steam and the tap water in it will evaporate. See Figure.
4. After a few minutes, when all of the tap water has evaporated, remove the dish and see if contains any solid residue.
5. Repeat the experiment with other water samples.



FIGURE

OBSERVATIONS

1. Why the given water sample is dried over a water bath ?
2. Do all waters have the same quantity of dissolved solids ?



EXPERIMENT 24 To prepare wood charcoalRELEVANCE TO TEXT 12 . 1PROCEDURE

1. Into a deep sand-bath place a number of small pieces of wood and completely cover with sand to exclude air.
2. Heat strongly for 30 minutes. Allow to cool then.
3. Separate wood residue from the sand.

OBSERVATIONS

1. How to account for the preparation of wood charcoal in absence of air ?

EXPERIMENT To show the absorption of colouring matter from solution by animal charcoal.

RELEVANCE AS ABOVEPROCEDURE

1. Half fill a 100-ml beaker with litmus solution.
2. Add a little (1 g) animal charcoal to the solution.
3. Boil for 10 minutes, and filter.
4. Examine the colour of the filtrate.
5. Repeat the experiment using diluted ink solution.

EXPERIMENT 25 To show that carbon is a reducing agent.

RELEVANCE TO TEXT 12 - 6

PROCEDURE

1. Scrape a small shallow hole in a charcoal block and place in it a little lead monoxide.
2. Light a bunsen flame and blow a gentle blast of flame, by means of a mouth blow-pipe, over the lead monoxide surface.
3. Continue blowing for 15 minutes.
4. Allow the residue to cool. Scrape it gently on a piece of paper. This is lead.

OBJECTIVES

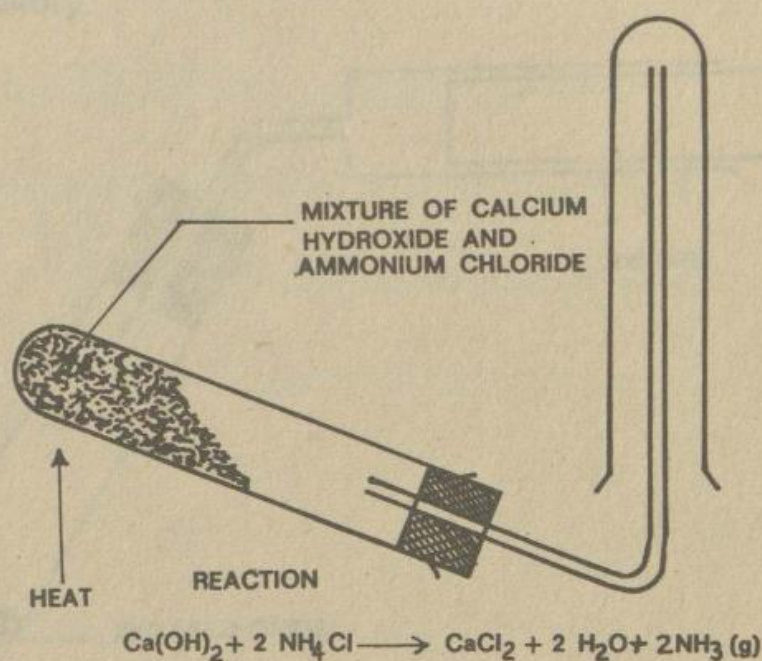
1. Why the residue is lead?
2. Write the reaction between lead oxide and carbon at high temperature.

EXPERIMENT 26 To prepare ammonia and investigate some of its properties.

RELEVANCE TO TEXT 13. 5

PROCEDURE

1. Mix together approximately equal quantities of solid calcium hydroxide $\text{Ca}(\text{OH})_2$ and ammonium chloride (NH_4Cl).
2. Grind the above mixture under folds of paper.
3. Half fill a hard test tube with the mixture.
4. Assemble the apparatus as shown in Figure.
5. Heat the test tube with a burner or lamp, and collect the gas given off as shown.

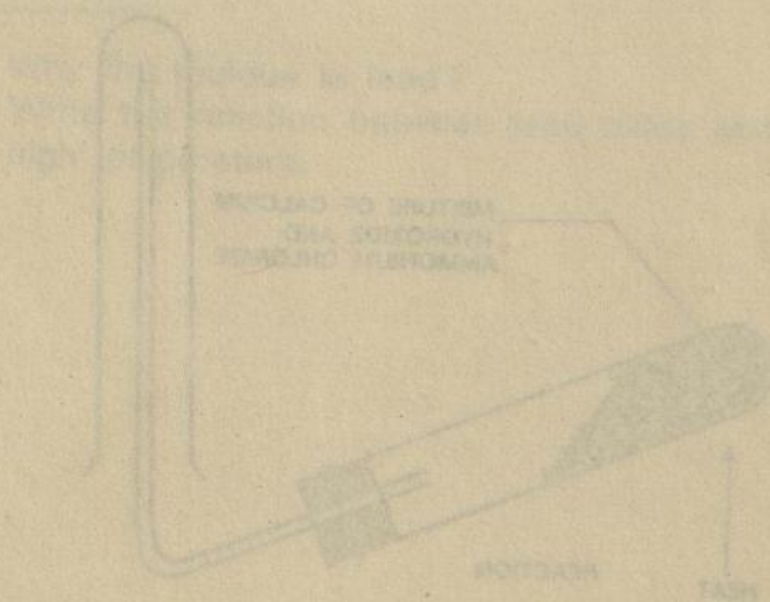


FIGURE

THE PREPARATION OF AMMONIA

OBSERVATIONS

1. Test the gas issuing from the delivery-tube with moist PH paper.
2. Note the appearance of the gas and its odour.
3. What happens when a stopper from a bottle of concentrated HCl is held near the delivery tube ?
4. Check the solubility of the gas in water.

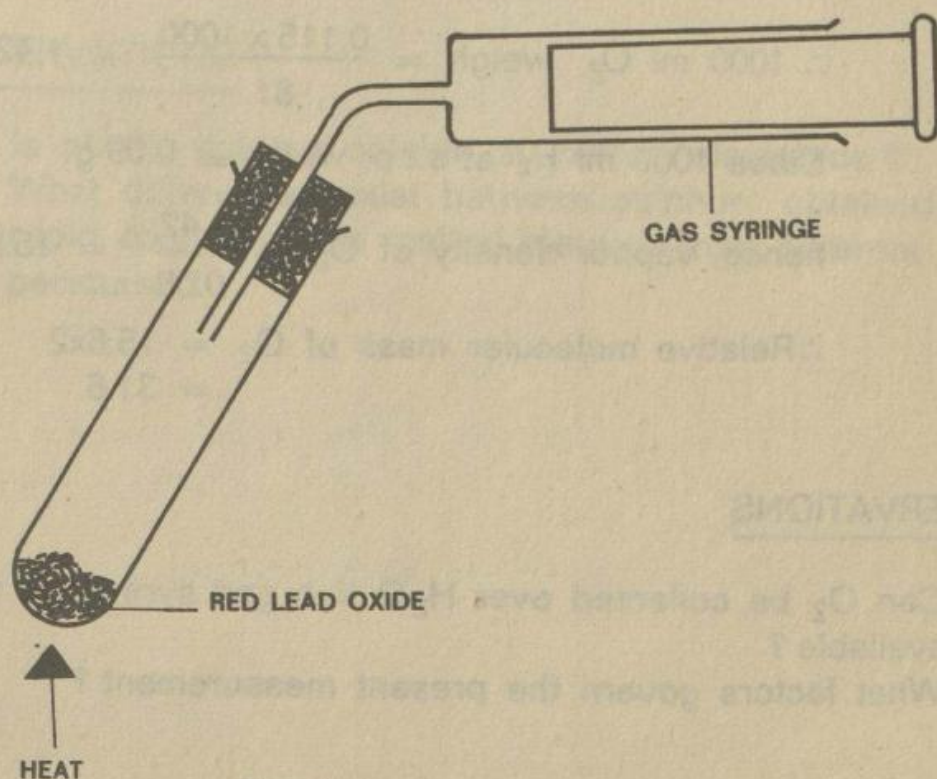


EXPERIMENT 27 To determine relative molecular mass of oxygen.

RELEVANCE TO TEXT 14 · 1

PROCEDURE

1. Weigh a hard glass test-tube containing about 4 g of dry red lead oxide.
2. Attach the tube to a gas syringe as shown in Figure.
3. Heat the tube gently, so that the O_2 produced is collected in the syringe.
4. Note the volume of the O_2 thus collected in syringe after the heating has been stopped.
5. Weigh the contents of the tube when cool.
6. Note room temperature and atmospheric pressure in laboratory.



FIGURE

SPECIMEN CALCULATION

Mass of tube + red lead oxide before heating = 17.397g

Mass of tube + contents after heating = 17.282g

Volume of O_2 collected = 84 cm^3

Temperature = 288 K°

Pressure = 750 mm Hg

Mass of O_2 = $17.397 - 17.282 = 0.115 \text{ g}$

Volume of O_2 at s.t.p. = $84 \times \frac{273}{288} \times \frac{750}{760} \text{ ml}$
 $= 81 \text{ ml}$

$\therefore 1000 \text{ ml } O_2 \text{ weigh} = \frac{0.115 \times 1000}{81} = 1.42 \text{ g}$

Since $1000 \text{ ml } H_2$ at s.t.p. weigh = 0.09 g

hence, vapour density of $O_2 = \frac{1.42}{0.09} = 15.8$

$\therefore$ Relative molecular mass of $O_2 = 15.8 \times 2$
 $= 31.6$

OBSERVATIONS

1. Can O_2 be collected over H_2O if a gas syringe is not available?
2. What factors govern the present measurement?

EXPERIMENT 28 To observe the effect of cooling on boiling sulphur suddenly

RELEVANCE TO TEXT 14 · 6

PROCEDURE

1. Half fill a test-tube with small pieces of roll sulphur.
2. Using a paper test-tube holder, heat the test-tube until sulphur melts.
3. Heat strongly until sulphur boils.
4. Carefully pour the boiling sulphur into a beaker of cold water.
5. Take the sulphur out of water and examine it. It is called plastic or gamma (γ) - sulphur.

OBSERVATIONS

1. Is plastic sulphur soluble in carbon disulphide?
2. What differences exist between sulphur obtained by rapid cooling of the melted element from different temperatures?

EXPERIMENT 29 To determine whether organic compounds contain carbon and hydrogen.

RELEVANCE TO TEXT 17 · 2, 17 · 3

PROCEDURE

1. Take a hard glass test tube and fill it to a depth of about half an inch with a mixture of little sugar and three times its amount of dried cupric oxide powder.
2. Holding the tube horizontally, heat the mixture strongly.
3. Pass any gas evolved through lime water. What is observed?
4. Note any condensate on the cooler parts of the tube. Add a little anhydrous CuSO_4 to the condensate on cooler parts of the tube. What is observed?
5. When the tube is cold, take out the residue on to a piece of paper and examine it. What has happened to cupric oxide?
6. Repeat the experiment with starch and sawdust.

OBSERVATIONS

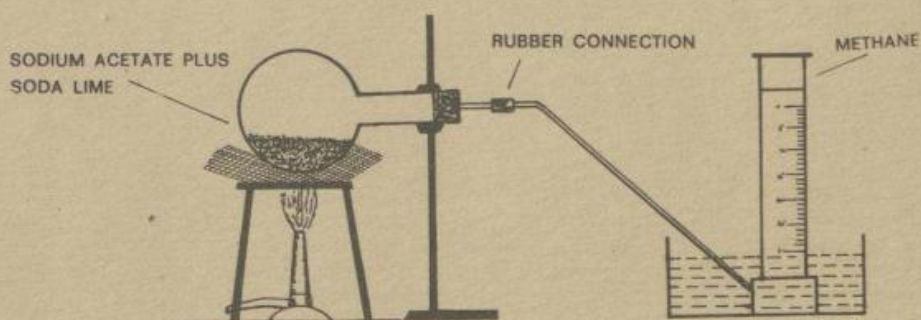
1. Copper oxide oxidizes all organic compounds when heated with them. Write down the reaction for the present experiment.
2. What is the evidence that organic compounds contain carbon and hydrogen?

EXPERIMENT 30 To prepare methane and study some of its properties

RELEVANCE TO TEXT 17 · 6

PROCEDURE

1. Fit up the apparatus shown in Figure.
2. Using a pestle and mortar, grind 21g of anhydrous sodium acetate with 42 g of sodium hydroxide.
3. Put the above mixture in a 200-ml flask filled with a bung carrying a delivery tube through a rubber tube.
4. Start heating gently, and then strongly as the gas starts bubbling up in the collecting jar. Remove the gas filled jar.
5. Collect 3 or 4 gas filled jars in the same way.
6. Test the gas jars for combustion of methane.
7. Note the colour/odour of the gas.
8. Test the gas for its solubility in water.
9. Test the gas for its acidic/basic character.



FIGURE

OBSERVATIONS

1. What precautions are necessary for methane production?
2. Write Chemical Equation for methane production.

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